### Resource Manager

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### CHAPTER 5 RESOURCE MANAGER

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**Key to National Science Content Standards:** UCP = Unifying Concepts and Processes, A = Science as Inquiry, B = Physical Science, C = Life Science, D = Earth and Space Sciences, E = Science and Technology, F = Science in Personal and Social Perspectives, G = History and Nature of Science

Refer to pages 4T–5T of the Teacher Guide for an explanation of the National Science Content Standards correlations.
Resource Manager

Materials List

**ChemLab** (pages 142–143)
- 40-W tubular light bulb, light socket with power cord, spectrum tubes (hydrogen, neon, and mercury), spectrum tube power supplies (3), Flinn C-Spectra®, diffraction grating, colored pencils, food coloring (red, green, blue, and yellow), 275-mL polystyrene culture flasks (4), book, water

**Discovery Lab** (page 117)
- wrapped box containing small object

**MiniLab** (page 125)
- Bunsen burner, cotton swabs (6), distilled water, lithium chloride, sodium chloride, potassium chloride, calcium chloride, strontium chloride, unknown

**Demonstration** (pages 136–137)
- spectrum tubes (hydrogen and neon), spectrum tube power supply, Flinn C-Spectra®, diffraction grating, colored pencils or chalk

Preparation of Solutions

For a review of solution preparation, see page 46T of the Teacher Guide.

There are no solutions to be prepared for the activities in this chapter.

Assessment Resources

- Chapter Assessment, pp. 25–30
- MindJogger Videoquizzes
- Alternate Assessment in the Science Classroom
- TestCheck Software
- Solutions Manual, Chapter 5
- Supplemental Problems, Chapter 5
- Performance Assessment in the Science Classroom
- Chemistry Interactive CD-ROM, Chapter 5 quiz

Additional Resources

- Spanish Resources
- Guided Reading Audio Program, Chapter 5
- Cooperative Learning in the Science Classroom
- Lab and Safety Skills in the Science Classroom
- Lesson Plans
- Block Scheduling Lesson Plans
- Texas Lesson Plans
- Texas Block Scheduling Lesson Plans
The following multimedia for this chapter are available from Glencoe.

**VIDEOTAPE/DVD**
MindJogger Videoquizzes, Chapter 5

**VIDEODISC**
Cosmic Chemistry
*Greenhouse Effect, Movie
Albert Einstein, Still
Niels Bohr, Still
Atomic Theories, Movie
Louis-Victor de Broglie, Still
Bohr-de Broglie Hydrogen Orbits, Still

**CD-ROM**
Chemistry: Matter and Change
*Flame Test, Video
The Aurora, Video
Atomic Emissions, Video
Electrons and Energy Levels, Animation
Bohr-de Broglie Hydrogen Orbits, Still

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**Multiple Learning Styles**

Look for the following icons for strategies that emphasize different learning modalities.

- **Kinesthetic**
  - Building a Model, p. 123; Meeting Individual Needs, pp. 127, 139; Quick Demo, p. 131

- **Visual-Spatial**
  - Portfolio, p. 118; Chemistry Journal, p. 133; Reteach, p. 141

- **Logical-Mathematical**
  - Meeting Individual Needs, p. 128

- **Linguistic**
  - Chemistry Journal, pp. 119, 122, 140; Portfolio, p. 133

- **Intrapersonal**
  - English Language Learners, p. 121; Enrichment, p. 122; Meeting Individual Needs, p. 131; Chemistry Journal, p. 129

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**Key to Teaching Strategies**

- **L1** Level 1 activities should be appropriate for students with learning difficulties.
- **L2** Level 2 activities should be within the ability range of all students.
- **L3** Level 3 activities are designed for above-average students.
- **ELL** ELL activities should be within the ability range of English Language Learners.
- **COOP LEARN** Cooperative Learning activities are designed for small group work.
- **P** These strategies represent student products that can be placed into a best-work portfolio.
- **P** These strategies are useful in a block scheduling format.

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**Assessment Planner**

- **Portfolio Assessment**
  - Portfolio, TWE, pp. 118, 120, 133, 145
  - Assessment, p. 139

- **Performance Assessment**
  - Assessment, TWE, pp. 122, 128
  - MiniLab, SE, p. 125
  - ChemLab, SE, pp. 142-143
  - Discovery Lab, SE, p. 117

- **Knowledge Assessment**
  - Assessment, TWE, pp. 126, 134
  - Section Assessment, SE, pp. 126, 134, 141
  - Chapter Assessment, SE, pp. 146-149

- **Skill Assessment**
  - Assessment, TWE, p. 141
  - Problem-Solving Lab, TWE, p. 130
  - ChemLab, TWE, p. 143
  - Demonstration, TWE, p. 137
Choosing to Previous Knowledge

Have students review the following concepts before studying this chapter.
Chapter 4: atomic structure

Using the Photo

Point out that the vivid colors of light given off by fireworks are of different origin than colors produced by colored light bulbs or filters. Explain that energy transitions within atoms cause the distinctive colors—something that students will learn more about in the chapter.

Chapter Themes

The following themes from the National Science Education Standards are covered in this chapter. Refer to page 4T of the Teacher Guide for an explanation of the correlations.
Systems, order, and organization (UCP.1)
Evidence, models, and explanation (UCP.2)
Form and function (UCP.5)

Why It’s Important

Why are some fireworks red, some white, and others blue? The key to understanding the chemical behavior of fireworks, and all matter, lies in understanding how electrons are arranged in atoms of each element.

VISIT THE CHEMISTRY WEBSITE AT

science.glencoe.com TO FIND LINKS ABOUT ELECTRONS AND ATOMIC STRUCTURE.

The colorful display from fireworks is due to changes in the electron configurations of atoms.

DISCOVERY LAB

Purpose

Students will make observations using all the senses except sight.

Safety and Disposal

Keep boxes for use next year.

Teaching Strategies

• Try to use objects in the box that are simple, but challenging.
• When students are through with the lab, you may want to identify the objects, or, to demonstrate that chemists can’t always see what they are looking for, you may want to leave the object’s identity a mystery!

Expected Results

Results will vary. Students should try to use senses other than sight to determine the relative size, mass, shape, and number of objects.

Analysis

Answers will vary. Students will determine that observations typically rely heavily upon sight, although touch and hearing are somewhat useful.
1 Focus

Focus Transparency

Before presenting the lesson, display Section Focus Transparency 17 on the overhead projector. Have students answer the accompanying questions using Section Focus Transparency Master 17. 

2 Teach

Concept Development

Explain that the concept that matter is made up of atoms is useful in many ways. For example, the fact that water contains two atoms of hydrogen for every atom of oxygen explains why the masses of the two elements are always in the same proportion in the compound. Point out, however, that something well beyond this concept must account for the vastly different chemical behaviors of hydrogen, oxygen, and the other chemical elements.

Light and Quantized Energy

Although three subatomic particles had been discovered by the early-1900s, the quest to understand the atom and its structure had really just begun. That quest continues in this chapter, as scientists pursued an understanding of how electrons were arranged within atoms. Perform the DISCOVERY LAB on this page to better understand the difficulties scientists faced in researching the unseen atom.

The Nuclear Atom and Unanswered Questions

As you learned in Chapter 4, Rutherford proposed that all of an atom’s positive charge and virtually all of its mass are concentrated in a nucleus that is surrounded by fast-moving electrons. Although his nuclear model was a major scientific development, it lacked detail about how electrons occupy the space surrounding the nucleus. In this chapter, you will learn how electrons are arranged in an atom and how that arrangement plays a role in chemical behavior.

Many scientists in the early twentieth century found Rutherford’s nuclear atomic model to be fundamentally incomplete. To physicists, the model did not explain how the atom’s electrons are arranged in the space around the nucleus. Nor did it address the question of why the negatively charged electrons are not pulled into the atom’s positively charged nucleus. Chemists found Rutherford’s nuclear model lacking because it did not begin to account for the differences in chemical behavior among the various elements.

Objectives

- Compare the wave and particle models of light.
- Define a quantum of energy and explain how it is related to an energy change of matter.
- Contrast continuous electromagnetic spectra and atomic emission spectra.

Vocabulary

electromagnetic radiation 
frequency 
amplitude 
emission spectrum 
quanta 
Planck’s constant 
photoelectric effect 
photon 
atomic emission spectrum

Analysis

How were you able to determine things such as size, shape, number, and composition of the object in the box? What senses did you use to make your observations? Why is it hard to figure out what type of object is in the box without actually seeing it?

Materials

a wrapped box from your instructor

What’s Inside?

It’s your birthday, and there are many wrapped presents for you to open. Much of the fun is trying to figure out what’s inside the package before you open it. In trying to determine the structure of the atom, chemists had a similar experience. How good are your skills of observation and deduction?

Procedure

1. Obtain a wrapped box from your instructor.
2. Using as many observation methods as you can, and without unwrapping or opening the box, try to figure out what the object inside the box is.
3. Record the observations you make throughout this discovery process.

Analysis

How were you able to determine things such as size, shape, number, and composition of the object in the box? What senses did you use to make your observations? Why is it hard to figure out what type of object is in the box without actually seeing it?
Quick Demo

Demonstrate that unlike charges attract using the following materials, which can probably be obtained from your school’s physics department: a hard rubber rod; either a piece of cat hide with the fur attached or a piece of wool; a glass rod; a piece of synthetic fabric, such as nylon. Use a Y-shaped piece of string to suspend the rubber rod horizontally from a support.

Impart a negative charge to the rod by rubbing it with fur. Then, impart a positive charge to the glass rod by rubbing it with synthetic fabric. When you bring the glass rod close to the suspended rubber rod, the rubber rod will move toward the glass rod. Explain that the rods’ unlike charges cause the attraction. Point out that an atom’s positively charged nucleus exerts the same type of electrostatic attraction for its negatively charged electrons.

Figure 5-1

- Chlorine gas, shown here reacting vigorously with steel wool, reacts with many other atoms as well.
- Argon gas fills the interior of this incandescent bulb. The nonreactive argon prevents the hot filament from oxidizing, thus extending the life of the bulb.
- Solid potassium metal is submerged in oil to prevent it from reacting with air or water.

For example, consider the elements chlorine, argon, and potassium, which are found in consecutive order on the periodic table but have very different chemical behaviors. Atoms of chlorine, a yellow-green gas at room temperature, react readily with atoms of many other elements. Figure 5-1a shows chlorine atoms reacting with steel wool. The interaction of highly reactive chlorine atoms with the large surface area provided by the steel results in a vigorous reaction. Argon, which is used in the incandescent bulb shown in Figure 5-1b, is also a gas. Argon, however, is so unreactive that it is considered a noble gas. Potassium is a reactive metal at room temperature. In fact, as you can see in Figure 5-1c, because potassium is so reactive, it must be stored under kerosene or oil to prevent its atoms from reacting with the oxygen and water in the air. Rutherford’s nuclear atomic model could not explain why atoms of these elements behave the way they do.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They had observed that certain elements emitted visible light when heated in a flame. Analysis of the emitted light revealed that an element’s chemical behavior is related to the arrangement of the electrons in its atoms. In order for you to better understand this relationship and the nature of atomic structure, it will be helpful for you to first understand the nature of light.

Wave Nature of Light

Electromagnetic radiation is a form of energy that exhibits wavelike behavior as it travels through space. Visible light is a type of electromagnetic radiation. Other examples of electromagnetic radiation include visible light from the sun, microwaves that warm and cook your food, X rays that doctors and dentists use to examine bones and teeth, and waves that carry radio and television programs to your home.

All waves can be described by several characteristics, a few of which you may be familiar with from everyday experience. Figure 5-2a shows a standing wave created by rhythmically moving the free end of a spring toy. Figure 5-2b illustrates several primary characteristics of all waves: wavelength, frequency, amplitude, and speed. Wavelength (represented by λ, the Greek letter lambda) is the shortest distance between equivalent points on a continuous wave. For example, in Figure 5-2b the wavelength is measured from crest to crest or from trough to trough. Wavelength is usually expressed in meters, centimeters, or nanometers (1 nm = 1 × 10⁻⁹ m). Frequency (represented by ν, the Greek letter nu) is the number of waves that pass a given point in one second. For visible light, frequency is approximately 6 × 10¹⁵ Hz.
point per second. One hertz (Hz), the SI unit of frequency, equals one wave per second. In calculations, frequency is expressed with units of “waves per second,” $\frac{1}{s}$ or (s$^{-1}$), where the term “waves” is understood. For example,

$$652 \text{ Hz} = 652 \text{ waves/second} = \frac{652}{s} = 652 \text{ s}^{-1}$$

The amplitude of a wave is the wave’s height from the origin to a crest, or from the origin to a trough. To learn how lightwaves are able to form powerful laser beams, read the How It Works at the end of this chapter.

All electromagnetic waves, including visible light, travel at a speed of $3.00 \times 10^8$ m/s in a vacuum. Because the speed of light is such an important and universal value, it is given its own symbol, $c$. The speed of light is the product of its wavelength ($\lambda$) and its frequency ($\nu$).

$$c = \lambda \nu$$

Although the speed of all electromagnetic waves is the same, waves may have different wavelengths and frequencies. As you can see from the equation above, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the red and violet light waves illustrated in Figure 5-3. Although both waves travel at the speed of light, you can see that red light has a longer wavelength and lower frequency than violet light.

Sunlight, which is one example of what is called white light, contains a continuous range of wavelengths and frequencies. Sunlight passing through a prism

**Figure 5-2**

The primary characteristics of waves are wavelength, frequency, amplitude, and speed. What is the wavelength of the wave in centimeters?

**Figure 5-3**

- The inverse relationship between wavelength and frequency of electromagnetic waves can be seen in these red and violet waves. As wavelength increases, frequency decreases. Wavelength and frequency do not affect the amplitude of a wave. Which wave has the larger amplitude?
- The red wave has a larger amplitude.

---

**Quick Demo**

Project the beam from a high-intensity projector into the side of a large beaker of water. Darken the room and adjust the arrangement so students can see the visible portion of the electromagnetic spectrum on a wall or screen. Explain that reflection and refraction separate the component colors of white light from the projector as they pass through the beaker and water. Point out that rainbows are formed in much the same manner when the colors in sunlight separate as they are reflected and refracted by raindrops.

---

**Figure Caption Questions**

**Figure 5-2** What is the wavelength of the wave in centimeters?

- ~4.5 cm

**Figure 5-3** Which wave has the larger amplitude? The red wave has a larger amplitude.

---

**Resource Manager**

**Math Skills Transparency 5 and Master L2 ELL**

**CHEMISTRY JOURNAL**

**Frequencies and Daily Living**

**Linguistic** In order to reinforce the concept of frequency, have students think of and describe at least five phenomena they encounter that recur or occur at given frequencies in their daily lives. Have them describe these phenomena in their journals and, when possible, quantify the frequencies.

---

**Chemistry TEKS**

Pages 118–119
3(C), 3(E)
Figure 5-5
The electromagnetic spectrum includes a wide range of wavelengths (and frequencies). Energy of the radiation increases with increasing frequency. Which types of waves or rays have the highest energy?

Visible light

Wavelengths ($\lambda$) in meters

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<tr>
<th>$3 \times 10^{-3}$</th>
<th>$3 \times 10^{-2}$</th>
<th>3</th>
<th>$3 \times 10^{-4}$</th>
<th>$3 \times 10^{-6}$</th>
<th>$3 \times 10^{-8}$</th>
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<tr>
<td>Radio</td>
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<td>Ultraviolet</td>
<td>Gamma rays</td>
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<tr>
<td>AM, TV, FM</td>
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Frequency ($\nu$) in hertz

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<tr>
<th>$10^4$</th>
<th>$10^5$</th>
<th>$10^6$</th>
<th>$10^7$</th>
<th>$10^8$</th>
<th>$10^9$</th>
<th>$10^{10}$</th>
<th>$10^{11}$</th>
<th>$10^{12}$</th>
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<td>Radio</td>
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<td>Gamma rays</td>
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Energy increases

Electromagnetic Spectrum

120 Chapter 5 Electrons in Atoms

Quick Demo
Borrow a Slinky from the physics department and attach it securely to an object on one side of the room. Demonstrate wave characteristics—wavelength, frequency, and energy—by generating standing waves. Start with a half wave, showing the longest wavelength, lowest frequency, and least energy. Work up to two or two and one-half standing waves. It will be obvious that you must use more energy as the number of standing waves increases. With each increase in the number of waves, ask students what is happening to frequency and wavelength, and how energy is changing. Frequency is increasing, wavelength is decreasing, and energy is increasing.

Figure Caption Question
Figure 5-5 Which types of waves or rays have the highest energy? gamma rays and X rays

Electromagnetic Waves and Uses
Have students research and discuss the many ways humans use electromagnetic waves to transmit information and carry energy from place to place. Have them write up their findings in their portfolios.

Resource Manager
Teaching Transparency 15 and Master L2 ELL
5.1 Light and Quantized Energy

Because all electromagnetic waves travel at the same speed, you can use the formula \( c = \lambda \nu \) to calculate the wavelength or frequency of any wave. Example Problem 5-1 shows how this is done.

**EXAMPLE PROBLEM 5-1**

**Calculating Wavelength of an EM Wave**

Microwaves are used to transmit information. What is the wavelength of a microwave having a frequency of \( 3.44 \times 10^9 \) Hz?

1. **Analyze the Problem**
   
   You are given the frequency of a microwave. You also know that because microwaves are part of the electromagnetic spectrum, their speed, frequency, and wavelength are related by the formula \( c = \lambda \nu \). The value of \( c \) is a known constant. First, solve the equation for wavelength, then substitute the known values and solve.

   Known Unknown
   
   \( \nu = 3.44 \times 10^9 \) Hz \( \lambda = ? \) m
   \( c = 3.00 \times 10^8 \) m/s

2. **Solve for the Unknown**

   Solve the equation relating the speed, frequency, and wavelength of an electromagnetic wave for wavelength (\( \lambda \)).

   \( c = \lambda \nu \)
   
   \( \lambda = \frac{c}{\nu} \)

   Substitute \( c \) and the microwave’s frequency, \( \nu \), into the equation. Note that hertz is equivalent to \( 1/\)s or \( s^{-1} \).

   \( \lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ s}^{-1}} \)

   Divide the values to determine wavelength, \( \lambda \), and cancel units as required.

   \( \lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ s}^{-1}} = 8.72 \times 10^{-2} \) m

3. **Evaluate the Answer**

   The answer is correctly expressed in a unit of wavelength (m). Both of the known values in the problem are expressed with three significant figures, so the answer should have three significant figures, which it does. The value for the wavelength is within the wavelength range for microwaves shown in Figure 5-5.

**PRACTICE PROBLEMS**

1. What is the frequency of green light, which has a wavelength of \( 4.90 \times 10^{-7} \) m?
2. An X ray has a wavelength of \( 1.15 \times 10^{-10} \) m. What is its frequency?
3. What is the speed of an electromagnetic wave that has a frequency of \( 7.8 \times 10^6 \) Hz?
4. A popular radio station broadcasts with a frequency of 94.7 MHz. What is the wavelength of the broadcast? (1 MHz = \( 10^6 \) Hz)

**Reinforcement**

When the people in a stadium make a “wave,” the wave travels around the stadium as individual persons move their bodies and arms up and down. Point out, however, that each person transmitting the wave remains in the same place.

**Math in Chemistry**

Explain that when two quantities are related mathematically in such a way that the increase in one quantity is proportional to the decrease in the other quantity, the two quantities are said to be inversely proportional. Point out that the relationship \( c = \lambda \nu \) is valid only if the quantities \( \lambda \) and \( \nu \) are inversely related.

**English Language Learners**

Intrapersonal Have English language learners look up and then explain the meanings of several key English words used in this section: radiation, spectrum, constant, effect, emission, quantum. Then ask them to use the words in a paragraph about waves.
While considering light as a wave does explain much of its everyday behavior, it fails to adequately describe important aspects of light’s interactions with matter. The wave model of light cannot explain why heated objects emit only certain frequencies of light at a given temperature, or why some metals emit electrons when colored light of a specific frequency shines on them. Obviously, a totally new model or a revision of the current model of light was needed to address these phenomena.

The quantum concept

The glowing light emitted by the hot objects shown in Figure 5-6 are examples of a phenomenon you have certainly seen. Iron provides another example of the phenomenon. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and appears bluish in color at even higher temperatures. As you will learn in greater detail later on in this course, the temperature of an object is a measure of the average kinetic energy of its particles. As the iron gets hotter it possesses a greater amount of energy, and emits different colors of light. These different colors correspond to different frequencies and wavelengths. The wave model could not explain the emission of these different wavelengths of light at different temperatures. In 1900, the German physicist Max Planck (1858–1947) began searching for an explanation as he studied the light emitted from heated objects. His study of the phenomenon led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. That is, a quantum is the minimum amount of energy that can be gained or lost by an atom.

Planck and other physicists of the time thought the concept of quantized energy was revolutionary—and some found it disturbing. Prior experience had led scientists to believe that energy could be absorbed and emitted in continually varying quantities, with no minimum limit to the amount. For example, think about heating a cup of water in a microwave oven. It seems that you can add any amount of thermal energy to the water by regulating the power and duration of the microwaves. Actually, the water’s temperature increases in infinitesimal steps as its molecules absorb quanta of energy. Because these steps are so small, the temperature seems to rise in a continuous, rather than a stepwise, manner.

The glowing objects shown in Figure 5-6 are emitting light, which is a form of energy. Planck proposed that this emitted light energy was quantized. The glowing light emitted by the hot objects shown in Figure 5-6 are examples of a phenomenon you have certainly seen. Iron provides another example of the phenomenon. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and appears bluish in color at even higher temperatures. As you will learn in greater detail later on in this course, the temperature of an object is a measure of the average kinetic energy of its particles. As the iron gets hotter it possesses a greater amount of energy, and emits different colors of light. These different colors correspond to different frequencies and wavelengths. The wave model could not explain the emission of these different wavelengths of light at different temperatures. In 1900, the German physicist Max Planck (1858–1947) began searching for an explanation as he studied the light emitted from heated objects. His study of the phenomenon led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. That is, a quantum is the minimum amount of energy that can be gained or lost by an atom.

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The glowing objects shown in Figure 5-6 are emitting light, which is a form of energy. Planck proposed that this emitted light energy was quantized.
He then went further and demonstrated mathematically that the energy of a quantum is related to the frequency of the emitted radiation by the equation

\[ E_{\text{quantum}} = h \nu \]

where \( E \) is energy, \( h \) is Planck's constant, and \( \nu \) is frequency. Planck's constant has a value of \( 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \), where \( \text{J} \) is the symbol for the joule, the SI unit of energy. Looking at the equation, you can see that the energy of radiation increases as the radiation's frequency, \( \nu \), increases. This equation explains why the violet light in Figure 5-3 has greater energy than the red light.

According to Planck's theory, for a given frequency, \( \nu \), matter can emit or absorb energy only in whole-number multiples of \( h \nu \); that is, \( 1h\nu, 2h\nu, 3h\nu \), and so on. A useful analogy for this concept is that of a child building a wall of wooden blocks. The child can add to or take away height from the wall only in increments of a whole number of blocks. Partial blocks are not possible. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

**The photoelectric effect** Scientists knew that the wave model (still very popular in spite of Planck's proposal) could not explain a phenomenon called the photoelectric effect. In the photoelectric effect, electrons, called photoelectrons, are emitted from a metal's surface when light of a certain frequency shines on the surface, as shown in Figure 5-7. Perhaps you've taken advantage of the photoelectric effect by using a calculator, such as the one shown in Figure 5-8, that is powered by photoelectric cells. Photoelectric cells in these and many other devices convert the energy of incident light into electrical energy.

The mystery of the photoelectric effect concerns the frequency, and therefore color, of the incident light. The wave model predicts that given enough time, even low-energy, low-frequency light would accumulate and supply enough energy to eject photoelectrons from a metal. However, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intense or how long it shines, light with a frequency less than \( 1.14 \times 10^{15} \text{ Hz} \) does not eject photoelectrons from silver. But even dim light having a frequency equal to or greater than \( 1.14 \times 10^{15} \text{ Hz} \) causes the ejection of photoelectrons from silver.

In explaining the photoelectric effect, Albert Einstein proposed in 1905 that electromagnetic radiation has both wavelike and particulate-like natures. That is, while a beam of light has many wavelike characteristics, it also can be thought of as a stream of tiny particles, or bundles of energy, called photons. Thus, a photon is a particle of electromagnetic radiation with no mass that carries a quantum of energy.

**Concept Development**

Explain to students that they might think of the light emitted by an atom as a "window into the atom." Explain further that the chemical behaviors of the elements are related not to the number of subatomic particles in their atoms, but to the arrangement of electrons within their atoms.

**Building a Model**

**Kinesthetic** Have student groups build a setup that models the photoelectric effect. For example, the setup might show that impacting small magnets attached to a heavy iron object with lightweight and low-energy objects such as marshmallows will not displace the magnets. Then, the setup could show that heavier objects with greater energy displace the magnets. Have students draw the analogy between the marshmallows and low-energy photons and between the heavier objects and high-energy photons.

**Internet Address Book**

Note Internet addresses that you find useful in the space below for quick reference.

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Chemistry TEKS

Pages 120–123
3(C), 3(E), 5(A)
Extending Planck’s idea of quantized energy, Einstein calculated that a photon’s energy depends on its frequency.

\[ E_{\text{photon}} = h \nu \]

Further, Einstein proposed that the energy of a photon of light must have a certain minimum, or threshold, value to cause the ejection of a photoelectron. That is, for the photoelectric effect to occur a photon must possess, at a minimum, the energy required to free an electron from an atom of the metal. According to this theory, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Although Einstein was able to explain the photoelectric effect by giving electromagnetic radiation particlelike properties, it’s important to note that a dual wave-particle model of light was required.
Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? The process illustrates another phenomenon that cannot be explained by the wave model of light. The light of the neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited and unstable atoms then release energy by emitting light. If the light emitted by the neon is passed through a glass prism, neon’s atomic emission spectrum is produced. The atomic emission spectrum of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. Neon’s atomic emission spectrum consists of several individual lines of color, not a continuous range of colors as seen in the visible spectrum.

Each element’s atomic emission spectrum is unique and can be used to determine if that element is part of an unknown compound. For example, when a platinum wire is dipped into a strontium nitrate solution and then inserted into a burner flame, the strontium atoms emit a characteristic red color. You can perform a series of flame tests yourself by doing the miniLAB below.

Figure 5-9 on the following page shows an illustration of the characteristic purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen’s emission spectrum responsible for producing the glow. Note how the line nature of hydrogen’s atomic emission spectrum differs from that of a continuous spectrum. To gain firsthand experience with types of line spectra, you can perform the CHEMLAB at the end of this chapter.

miniLAB

Flame Tests

Classifying When certain compounds are heated in a flame, they emit a distinctive color. The color of the emitted light can be used to identify the compound.

Materials Bunsen burner; cotton swabs (6); distilled water; crystals of lithium chloride, sodium chloride, potassium chloride, calcium chloride, strontium chloride, unknown

Procedure

1. Dip a cotton swab into the distilled water. Dip the moistened swab into the lithium chloride so that a few of the crystals stick to the cotton. Put the crystals on the swab into the flame of a Bunsen burner. Observe the color of the flame and record it in your data table.
2. Repeat step 1 for each of the metallic chlorides (sodium chloride, potassium chloride, calcium chloride, and strontium chloride). Be sure to record the color of each flame in your data table.
3. Obtain a sample of unknown crystals from your teacher. Repeat the procedure in step 1 using the unknown crystals. Record the color of the flame produced by the unknown crystals in your data table. Dispose of used cotton swabs as directed by your teacher.

Analysis

1. Each of the known compounds tested contains chlorine, yet each compound produced a flame of a different color. Explain why this occurred.
2. How is the atomic emission spectrum of an element related to these flame tests?
3. What is the identity of the unknown crystals? Explain how you know.

Flame Test Results

<table>
<thead>
<tr>
<th>Compound</th>
<th>Flame color</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium chloride</td>
<td>red</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>yellow</td>
</tr>
<tr>
<td>Potassium chloride</td>
<td>violet</td>
</tr>
<tr>
<td>Calcium chloride</td>
<td>red-orange</td>
</tr>
<tr>
<td>Strontium chloride</td>
<td>bright red</td>
</tr>
<tr>
<td>Unknown</td>
<td>depends on compound</td>
</tr>
</tbody>
</table>

Analysis

1. The colors are due primarily to electron transitions of the metal atoms. The colors are characteristic of lithium, sodium, potassium, calcium, and strontium.
2. The colors are a composite of each element’s visible spectrum.
3. Answers will vary depending on the identity of the unknown sample.

Assessment

Performance Have students look at the flame color spectra using a Flinn C-Spectra or a spectroscope and relate the spectra to the elements comprising each compound. Use the Performance Task Assessment List for Analyzing the Data in PASC, p. 27.

Chemistry TEKS

Pages 124–125
1(A), 2(B), 2(E), 3(C), 3(E), 5(A)
Reinforce the concept that red light has less energy than blue light. Explaining that you are making a solution of a fluorescent substance, prepare a solution of about 10 g fluorescein in 100 mL water in a 150 mL beaker. Turn out the room lights, and shine a flashlight’s beam through a transparent red cellophane sheet into the fluorescein solution. When you turn out the flashlight, the solution will not fluoresce. Then, repeat the process, but use a blue cellophane sheet rather than a red one. The solution may be flushed down the drain with water.

CHEMLAB

ChemLab 5, located at the end of the chapter, can be used at this point in the lesson.

Assessment

Knowledge Ask students to compare the wavelengths, frequencies, and energies of microwaves and X rays. Microwaves have longer wavelengths, lower frequencies, and lower energies than X rays.

Figure Caption Question

Figure 5-9 Which line has the highest energy? the violet line

Section 5.1 Assessment

1. List the characteristic properties of all waves. At what speed do electromagnetic waves travel in a vacuum?
2. Compare the wave and particle models of light. What phenomena can only be explained by the particle model?
3. What is a quantum of energy? Explain how quanta of energy are involved in the amount of energy matter gains and loses.
4. Explain the difference between the continuous spectrum of white light and the atomic emission spectrum of an element.

Figure 5-5 and your knowledge of light to match the numbered items on the right with the lettered items on the left. The numbered items may be used more than once or not at all.

a. longest wavelength 1. gamma rays
b. highest frequency 2. infrared waves
c. greatest energy 3. radio waves

d. speed, wavelength, frequency, and amplitude; EM waves travel at c. 

The wave model treats light as an electromagnetic wave. The particle model treats light as being comprised of photons. The wave model could not explain the photoelectric effect, the color of hot objects, and emission spectra.

A quantum is the minimum amount of energy that can be lost or gained by an atom. Matter loses or gains energy in multiples of the quantum.

A continuous spectrum contains all the visible colors; an atomic emission spectrum contains only specific colors.

Einstein proposed that electromagnetic radiation has a wave-particle nature, that the energy of a photon depends on the frequency of the radiation, and that the photon’s energy is given by the formula $E_{\text{photon}} = h\nu$.

11. Thinking Critically Explain how Einstein utilized Planck’s quantum concept in explaining the photoelectric effect.

12. Interpreting Scientific Illustrations Use Figure 5-5 and your knowledge of light to match the numbered items on the right with the lettered items on the left. The numbered items may be used more than once or not at all.

a. longest wavelength 1. gamma rays
b. highest frequency 2. infrared waves
c. greatest energy 3. radio waves

An atomic emission spectrum is characteristic of the element being examined and can be used to identify that element. The fact that only certain colors appear in an element’s atomic emission spectrum means that only certain specific frequencies of light are emitted. And because those emitted frequencies of light are related to energy by the formula $E_{\text{photon}} = h\nu$, it can be concluded that only photons having certain specific energies are emitted. This conclusion was not predicted by the laws of classical physics known at that time. Scientists found atomic emission spectra puzzling because they had expected to observe the emission of a continuous series of colors and energies as excited electrons lost energy and spiraled toward the nucleus. In the next section, you will learn about the continuing development of atomic models, and how one of those models was able to account for the frequencies of the light emitted by excited atoms.
You now know that the behavior of light can be explained only by a dual wave-particle model. Although this model was successful in accounting for several previously unexplainable phenomena, an understanding of the relationships among atomic structure, electrons, and atomic emission spectra still remained to be established.

**Bohr Model of the Atom**

Recall that hydrogen’s atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are elements’ atomic emission spectra discontinuous rather than continuous? Niels Bohr, a young Danish physicist working in Rutherford’s laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Impressively, Bohr’s model also correctly predicted the frequencies of the lines in hydrogen’s atomic emission spectrum.

**Energy states of hydrogen** Building on Planck’s and Einstein’s concepts of quantized energy (quantized means that only certain values are allowed), Bohr proposed that the hydrogen atom has only certain allowable energy states. The lowest allowable energy state of an atom is called its **ground state**. When an atom gains energy, it is said to be in an excited state. And although a hydrogen atom contains only a single electron, it is capable of having many different excited states.

Bohr went even further with his atomic model by relating the hydrogen atom’s energy states to the motion of the electron within the atom. Bohr suggested that the single electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron’s orbit, the lower the atom’s energy state, or energy level. Conversely, the larger the electron’s orbit, the higher the atom’s energy state, or energy level. Bohr assigned a quantum number, \( n \), to each orbit and even calculated the orbit’s radius. For the first orbit, the one closest to the nucleus, \( n = 1 \) and the orbit radius is 0.0529 nm; for the second orbit, \( n = 2 \) and the orbit radius is 0.212 nm; and so on. Additional information about Bohr’s description of hydrogen’s allowable orbits and energy levels is given in Table 5-1.

**Vocabulary**

- **ground state**
- **de Broglie equation**
- **Heisenberg uncertainty principle**
- **quantum mechanical model**
- **orbital**
- **energy sublevel**
- **principal energy level**
- **atomic orbital**
- **principal quantum number**

**Table 5-1** Bohr’s Description of the Hydrogen Atom

<table>
<thead>
<tr>
<th>Bohr atomic orbit</th>
<th>Quantum number</th>
<th>Orbit radius (nm)</th>
<th>Corresponding atomic energy level</th>
<th>Relative energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>First</td>
<td>( n = 1 )</td>
<td>0.0529</td>
<td>1</td>
<td>( E_1 )</td>
</tr>
<tr>
<td>Second</td>
<td>( n = 2 )</td>
<td>0.212</td>
<td>2 ( E_2 = 4E_1 )</td>
<td></td>
</tr>
<tr>
<td>Third</td>
<td>( n = 3 )</td>
<td>0.476</td>
<td>3 ( E_3 = 9E_1 )</td>
<td></td>
</tr>
<tr>
<td>Fourth</td>
<td>( n = 4 )</td>
<td>0.846</td>
<td>4 ( E_4 = 16E_1 )</td>
<td></td>
</tr>
<tr>
<td>Fifth</td>
<td>( n = 5 )</td>
<td>1.32</td>
<td>5 ( E_5 = 25E_1 )</td>
<td></td>
</tr>
<tr>
<td>Sixth</td>
<td>( n = 6 )</td>
<td>1.90</td>
<td>6 ( E_6 = 36E_1 )</td>
<td></td>
</tr>
<tr>
<td>Seventh</td>
<td>( n = 7 )</td>
<td>2.59</td>
<td>7 ( E_7 = 49E_1 )</td>
<td></td>
</tr>
</tbody>
</table>

**Objectives**

- **Compare** the Bohr and quantum mechanical models of the atom.
- **Explain** the impact of de Broglie’s wave-particle duality and the Heisenberg uncertainty principle on the modern view of electrons in atoms.
- **Identify** the relationships among a hydrogen atom’s energy levels, sublevels, and atomic orbitals.

**Learning Disabled**

- **Kinesthetic**: Demonstrate the electron transitions associated with energy-level changes. Tell students that a book on the floor represents an electron in an atom’s lowest-energy orbit. Raise the book to a higher energy level (their chair). Ask if energy is required. **Yes** Ask what happens when the book returns to the floor. **Energy is released**. Explain the analogy between the book’s energy levels and an electron’s transitions between atomic orbits. Point out that the energy needed to raise an electron to a higher-energy orbit is exactly the same as the energy released when the electron returns to its original orbit. **L1 ELL**
the following orbit transitions and the appropriate orbits to simulate them move the thumbtack between lowest energy state. Then, have emission spectrum: violet (6 → 2), blue-violet (5 → 2), blue-green (4 → 2), and red (3 → 2). Use the Performance Task Assessment List for Model in PASC, p. 51.

**Figure 5-10**

- When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon with a specific energy is emitted. Although hydrogen has spectral lines associated with higher energy levels, only the visible, ultraviolet, and infrared series of spectral lines are shown in this diagram. The relative energies of the electron transitions responsible for hydrogen’s four visible spectral lines are shown. Note how the energy levels become more closely spaced as \( n \) increases.

**An explanation of hydrogen’s line spectrum** Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when the electron is in the \( n = 1 \) orbit. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit such as the \( n = 2 \) orbit shown in Figure 5-10a. Such an electron transition raises the atom to an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit. As a result of this transition, the atom emits a photon corresponding to the difference between the energy levels associated with the two orbits.

\[
\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu
\]

Note that because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted. You might compare hydrogen’s seven atomic orbits to seven rungs on a ladder. A person can climb up or down the ladder only from rung to rung. Similarly, the hydrogen atom’s electron can move only from one allowable orbit to another, and therefore, can emit or absorb only certain amounts of energy.

The four electron transitions that account for visible lines in hydrogen’s atomic emission spectrum are shown in Figure 5-10b. For example, electrons dropping from the third orbit to the second orbit cause the red line. Note that electron transitions from higher-energy orbits to the second orbit account for all of hydrogen’s visible lines. This series of visible lines is called the Balmer series. Other electron transitions have been measured that are not visible, such as the Lyman series (ultraviolet) in which electrons drop into the \( n = 1 \) orbit and the Paschen series (infrared) in which electrons drop into the \( n = 3 \) orbit. Figure 5-10b also shows that unlike rungs on a ladder, the hydrogen atom’s energy levels are not evenly spaced. You will be able to see in greater detail how Bohr’s atomic model was able to account for hydrogen’s line spectrum by doing the problem-solving LAB later in this chapter.

Bohr’s model explained hydrogen’s observed spectral lines remarkably well. Unfortunately, however, the model failed to explain the spectrum of any other element. Moreover, Bohr’s model did not fully account for the chemical behavior of atoms. In fact, although Bohr’s idea of quantized energy levels laid the groundwork for atomic models to come, later experiments demonstrated that the Bohr model was fundamentally incorrect. The movements of electrons in atoms are not completely understood even now; however, substantial evidence indicates that electrons do not move around the nucleus in circular orbits.

**Gifted**

- **Logical-Mathematical** Ask gifted students to work through Bohr’s use of Newton’s second law \( F = ma \), Coulomb’s constant \( K \), and Bohr’s own concept of quantized angular momenta to derive the relationship \( r_n = \frac{\hbar^2 n^2}{4\pi^2 K a} \). Then, have them use the equation to calculate the radii of the hydrogen atom’s first four Bohr orbits.
The Quantum Mechanical Model of the Atom

Scientists in the mid-1920s, by then convinced that the Bohr atomic model was incorrect, formulated new and innovative explanations of how electrons are arranged in atoms. In 1924, a young French graduate student in physics named Louis de Broglie (1892–1987) proposed an idea that eventually accounted for the fixed energy levels of Bohr’s model.

Electrons as waves

De Broglie had been thinking that Bohr’s quantized electron orbits had characteristics similar to those of waves. For example, as Figure 5-11b shows, only multiples of half-wavelengths are possible on a plucked guitar string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in Figure 5-11c. He also reflected on the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question. If waves can have particle-like behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

Figure 5-11

A vibrating guitar string is constrained to vibrate between two fixed end points. The possible vibrations of the guitar string are limited to multiples of half-wavelengths. Thus, the “quantum” of the guitar string is one-half wavelength. The possible circular orbits of an electron are limited to whole numbers of complete wavelengths.

Gases for IR and UV

**Intrapersonal** Have students research the types of gases used to emit infrared and ultraviolet electromagnetic radiation. Have them summarize their findings in their chemistry journals.

**MEETING INDIVIDUAL NEEDS**

**English Language Learners**

Have English language learners look up and then explain the meanings of several key English words used in this section: state (as in ground state), uncertainty, principal, level (noun). Ask students to use each word in a sentence or a paragraph.

**Quick Demo**

Have a fan rotating at high speed when students enter the classroom so that they will not have seen the fan’s blades in a stopped position. As soon as the class period begins, ask them to describe the fan’s blades. They will be able to tell the blades’ approximate length and little else. Explain that scientists experience somewhat the same situation in trying to describe electrons in atoms. The electrons move about the nucleus and appear to fill the entire volume, yet occupy very little volume themselves. Explain that due to the motion of the electrons and certain limitations in our ability to view them (as described by Heisenberg’s uncertainty principle), we are unable to simultaneously describe exactly where the electrons are and where they are going.

**CHEMISTRY JOURNAL**

**Lab Manual, pp. 33–36**

**GASES**

**CHEMISTRY TEKS**

**Pages 128–129**

3(C), 3(E), 6(A)
# Problem-Solving Lab

## Purpose
Students will explore the relationship between electron orbit radii and energy states of the hydrogen atom. This relationship will then be used to explain the characteristics of spectroscopic series that result from electron transitions between orbits.

## Process Skills
Constructing models, using numbers, acquiring and analyzing information, drawing conclusions, applying concepts, predicting

## Teaching Strategies
- Ask students to explain how the force between two magnets depends on the separation distance, and then relate this to the electric force of attraction between an electron and proton. The magnetic force decreases with the cube of the distance. Because of this, magnets are useful in modeling the behavior of electric force, which decreases with the square of the distance.
- Ask students to describe the full range of the electromagnetic spectrum and how frequency, $\nu$, wavelength, $\lambda$, and energy, $E$, are related. $c = \lambda \nu$ ($c$ is the speed of light) and $E = h\nu$

## Thinking Critically
1. The largest radius is $r_1 = (2.9 \times 10^{-10} \text{ m})$. Thus, a scale of 1 cm = $1 \times 10^{-10} \text{ m}$ results in a graph about 26 cm in diameter, which will fit on the two sheets of paper. Energy increases with increasing radius.
2. See the Solutions Manual.
3. See the Solutions Manual.
4. greatest energy to least energy: Lyman, Balmer, Paschen, Brackett, and Pfund series
5. See the Solutions Manual.
6. Lyman series, ultra-violet; Balmer series, visible; Paschen series, near infrared; Brackett series, middle infrared; Pfund series, far infrared

## Analysis
Using the orbit radii equation, calculate hydrogen’s first seven electron orbit radii and then construct a scale model of those orbits. Use a compass and a metric ruler to draw your scale model on two sheets of paper that have been taped together. (Use caution when handling sharp objects.) Using the orbit energy equation, calculate the energy of each electron orbit and record the values on your model.

<table>
<thead>
<tr>
<th>Using Models</th>
<th>Thinking Critically</th>
</tr>
</thead>
<tbody>
<tr>
<td>Niels Bohr proposed that electrons must occupy specific, quantized energy levels in an atom. He derived the following equations for hydrogen’s electron orbit energies ($E_n$) and radii ($r_n$).</td>
<td>1. What scale did you use to plot the orbits? How is the energy of each orbit related to its radius?</td>
</tr>
<tr>
<td>$r_n = (0.529 \times 10^{-10} \text{ m})n^2$</td>
<td>2. Draw a set of arrows for electron jumps that end at each energy level (quantum number). For example, draw a set of arrows for all transitions that end at $n = 1$, a set of arrows for all transitions that end at $n = 2$, and so on, up to $n = 7$.</td>
</tr>
<tr>
<td>$E_n = -(2.18 \times 10^{-18} \text{ J})n^2$</td>
<td>3. Calculate the energy released for each of the jumps in step 2, and record the values on your model. The energy released is equal to the difference in the energies of each level.</td>
</tr>
<tr>
<td>Where $n =$ quantum number (1, 2, 3...).</td>
<td>4. Each set of arrows in step 2 represents a spectral emission series. Label five of the series, from greatest energy change to least energy change, as the Lyman, Balmer, Paschen, Brackett, and Pfund series.</td>
</tr>
</tbody>
</table>

## In-Depth Safety

**Science Safety**

- Characteristics
- Heating
- Placement
- Energy

**Step-by-Step Procedure**

1. Question: How was Bohr’s atomic model able to explain the line spectrum of hydrogen?

2. Purpose: Students will explore the relationship between electron orbit radii and energy states of the hydrogen atom. This relationship will then be used to explain the characteristics of spectroscopic series that result from electron transitions between orbits.

3. Process Skills: Constructing models, using numbers, acquiring and analyzing information, drawing conclusions, applying concepts, predicting

4. Teaching Strategies:
   - Ask students to explain how the force between two magnets depends on the separation distance, and then relate this to the electric force of attraction between an electron and proton. The magnetic force decreases with the cube of the distance. Because of this, magnets are useful in modeling the behavior of electric force, which decreases with the square of the distance.
   - Ask students to describe the full range of the electromagnetic spectrum and how frequency, $\nu$, wavelength, $\lambda$, and energy, $E$, are related. $c = \lambda \nu$ ($c$ is the speed of light) and $E = h\nu$

5. Thinking Critically:
   - The largest radius is $r_1 = (2.9 \times 10^{-10} \text{ m})$. Thus, a scale of 1 cm = $1 \times 10^{-10} \text{ m}$ results in a graph about 26 cm in diameter, which will fit on the two sheets of paper. Energy increases with increasing radius.
   - See the Solutions Manual.
   - See the Solutions Manual.
   - greatest energy to least energy: Lyman, Balmer, Paschen, Brackett, and Pfund series
   - See the Solutions Manual.
   - Lyman series, ultra-violet; Balmer series, visible; Paschen series, near infrared; Brackett series, middle infrared; Pfund series, far infrared

6. Analysis:
   - Using the orbit radii equation, calculate hydrogen’s first seven electron orbit radii and then construct a scale model of those orbits. Use a compass and a metric ruler to draw your scale model on two sheets of paper that have been taped together. (Use caution when handling sharp objects.) Using the orbit energy equation, calculate the energy of each electron orbit and record the values on your model.

7. In considering this question, de Broglie knew that if an electron has wave-like motion and is restricted to circular orbits of fixed radius, the electron is allowed only certain possible wavelengths, frequencies, and energies. Developing his idea, de Broglie derived an equation for the wavelength ($\lambda$) of a particle of mass ($m$) moving at velocity ($v$).

   \[ \lambda = \frac{h}{mv} \]

   The de Broglie equation predicts that all moving particles have wave characteristics. Why, then, you may be wondering, haven’t you noticed the wavelength of a fast-moving automobile? Using de Broglie’s equation provides an answer. An automobile moving at 25 m/s and having a mass of 910 kg has a wavelength of $2.9 \times 10^{-3}$ m—a wavelength far too small to be seen or detected, even with the most sensitive scientific instrument. By comparison, an electron moving at the same speed has the easily measured wavelength of $2.9 \times 10^{-5}$ m. Subsequent experiments have proven that electrons and other moving particles do indeed have wave characteristics.

   Step by step, scientists such as Rutherford, Bohr, and de Broglie had been unraveling the mysteries of the atom. However, a conclusion reached by the German theoretical physicist Werner Heisenberg (1901–1976), a contemporary of de Broglie, proved to have profound implications for atomic models.

**Assessment**

**Skill** Have the students extend the ideas presented here to make a prediction concerning the spectrum that would be emitted from hydrogen-like atoms, such as He$^+$ or Li$^{2+}$. Or, have them predict what would happen to the continuous spectrum of light if it passed through a cell containing hydrogen gas. **L2**
The Heisenberg Uncertainty Principle

Heisenberg’s concluded that it is impossible to make any measurement on an object without disturbing the object—at least a little. Imagine trying to locate a hovering, helium-filled balloon in a completely darkened room. When you wave your hand about, you’ll locate the balloon’s position when you touch it. However, when you touch the balloon, even gently, you transfer energy to it and change its position. Of course, you could also detect the balloon’s position by turning on a flashlight. Using this method, photons of light that reflect from the balloon reach your eyes and reveal the balloon’s location. Because the balloon is much more massive than the photons, the rebounding photons have virtually no effect on the balloon’s position.

Can photons of light help determine the position of an electron in an atom? As a thought experiment, imagine trying to determine the electron’s location by “bumping” it with a high-energy photon of electromagnetic radiation. Unfortunately, because such a photon has about the same energy as an electron, the interaction between the two particles changes both the wavelength of the photon and the position and velocity of the electron, as shown in Figure 5-12. In other words, the act of observing the electron produces a significant, unavoidable uncertainty in the position and motion of the electron. Heisenberg’s analysis of interactions such as those between photons and electrons led him to his historic conclusion. The Heisenberg uncertainty principle states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Although scientists of the time found Heisenberg’s principle difficult to accept, it has been proven to describe the fundamental limitations on what can be observed. How important is the Heisenberg uncertainty principle? The interaction of a photon with an object such as a helium-filled balloon has so little effect on the balloon that the uncertainty in its position is too small to measure. But that’s not the case with an electron moving at $6 \times 10^6 \text{ m/s}$ near an atomic nucleus. The uncertainty in the electron’s position is at least $10^{-3} \text{ m}$, about ten times greater than the diameter of the entire atom!

The Schrödinger wave equation In 1926, Austrian physicist Erwin Schrödinger (1887–1961) furthered the wave-particle theory proposed by de Broglie. Schrödinger derived an equation that treated the hydrogen atom’s electron as a wave. Remarkably, Schrödinger’s new model for the hydrogen atom seemed to apply equally well to atoms of other elements—an area in which Bohr’s model failed. The atomic model in which electrons are treated as waves is called the wave mechanical model of the atom. Like Bohr’s model, quantum mechanics is needed to accurately describe and explain atomic and subatomic behavior.
A photon striking an atom in an excited state stimulates it to make a transition to a lower-energy state and emit a second photon coherent with the first. Coherent means that the photons have the same associated wavelengths and are in phase (crest-to-crest and trough-to-trough). In a laser, photons from many atoms are reflected back and forth until they build to an intense, small beam—typically about 0.5 mm in diameter.

Medical lasers can be engineered to produce pulses of varying wavelength, intensity, and duration. For example, ophthalmologists can reshape corneas by removing tissue with 10-ns pulses from a 193-nm wavelength argon laser.

Because laser beams can be focused to such small diameters, they can be used for internal surgeries, destroying target tissue without adversely affecting surrounding tissue. And by channeling laser beams through optical fibers, doctors can perform surgeries in previously unreachable parts of the body. For example, bundles of optical fibers threaded through arteries can carry laser beams that destroy blockages.

**Enrichment**

Students may think the letters s, p, d, and f, which represent sublevels, arbitrary and perhaps mysterious. Explain that the letters originated from descriptions of spectral lines as sharp, principal, diffuse, and fundamental.

The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function. And most importantly, the wave function is related to the probability of finding the electron within a particular volume of space around the nucleus. Recall from your study of math that an event having a high probability is more likely to occur than one having a low probability.

What does the wave function predict about the electron’s location in an atom? A three-dimensional region around the nucleus called an atomic orbital describes the electron’s probable location. You can picture an atomic orbital as a fuzzy cloud in which the density of the cloud at a given point is proportional to the probability of finding the electron at that point. Figure 5-13a illustrates the probability map, or orbital, that describes the hydrogen electron in its lowest energy state. It might be helpful to think of the probability map as a time-exposure photograph of the electron moving around the nucleus, in which each dot represents the electron’s location at an instant in time. Because the dots are so numerous near the positive nucleus, they seem to form a dense cloud that is indicative of the electron’s most probable location. However, because the cloud has no definite boundary, it also is possible that the electron might be found at a considerable distance from the nucleus.

**Hydrogen’s Atomic Orbitals**

Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exactly defined size. To overcome the inherent uncertainty about the electron’s location, chemists arbitrarily draw an orbital’s surface to contain 90% of the electron’s total probability distribution. In other words, the electron spends 90% of the time within the volume defined by the surface, and 10% of the time somewhere outside the surface. The spherical surface shown in Figure 5-13b encloses 90% of the lowest-energy orbital of hydrogen.

Recall that the Bohr atomic model assigns quantum numbers to electron orbits. In a similar manner, the quantum mechanical model assigns principal quantum numbers \( n \) that indicate the relative sizes and energies of atomic orbitals. The quantum mechanical model limits an electron’s energy to certain values.

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The quantum mechanical model limits an electron’s energy to certain values. However, unlike Bohr’s model, the quantum mechanical model makes no attempt to describe the electron’s path around the nucleus.

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orbital becomes larger, the electron spends more time farther from the nucleus, and the atom’s energy level increases. Therefore, \( n \) specifies the atom’s major energy levels, called principal energy levels. An atom’s lowest principal energy level is assigned a principal quantum number of one. When the hydrogen atom’s single electron occupies an orbital with \( n = 1 \), the atom is in its ground state. Up to seven energy levels have been detected for the hydrogen atom, giving \( n \) values ranging from 1 to 7.

Principal energy levels contain energy sublevels. Principal energy level 1 consists of a single sublevel, principal energy level 2 consists of two sublevels, principal energy level 3 consists of three sublevels, and so on. To better understand the relationship between the atom’s energy levels and sublevels, picture the seats in a wedge-shaped section of a theater, as shown in Figure 5-14. As you move away from the stage, the rows become higher and contain more seats. Similarly, the number of energy sublevels in a principal energy level increases as \( n \) increases.

Sublevels are labeled s, p, d, or f according to the shapes of the atom’s orbitals. All s orbitals are spherical and all p orbitals are dumbbell-shaped; however, not all d or f orbitals have the same shape. Each orbital may contain at most two electrons. The single sublevel in principal energy level 1 consists of a spherical orbital called the 1s orbital. The two sublevels in principal energy level 2 are designated 2s and 2p. The 2s sublevel consists of the 2s orbital, which is spherical like the 1s orbital but larger in size. See Figure 5-15a. The 2p sublevel consists of three dumbbell-shaped p orbitals of equal energy designated 2p\(_x\), 2p\(_y\), and 2p\(_z\). The subscripts \( x \), \( y \), and \( z \) merely designate the orientations of p orbitals along the \( x \), \( y \), and \( z \) coordinate axes, as shown in Figure 5-15b.

Principal energy level 3 consists of three sublevels designated 3s, 3p, and 3d. Each d sublevel consists of five orbitals of equal energy. Four d orbitals have identical shapes but different orientations. However, the fifth, d\(_z^2\) orbital is shaped and oriented differently from the other four. The shapes and orientations of the five d orbitals are illustrated in Figure 5-16. The fourth principal energy level (\( n = 4 \)) contains a fourth sublevel, called the 4f sublevel, which consists of seven f orbitals of equal energy.

**Identifying Misconceptions**

Students may think that the hydrogen atom’s energy levels are evenly spaced.

**Uncover the Misconception**

Have students compare hydrogen’s energy levels shown in Figure 5-10b with the rungs on a ladder. Unlike the rungs on a ladder, hydrogen’s energy levels are not evenly spaced.

**Demonstrate the Concept**

Have students calculate and compare the ratios \( E_n/E_{n-1} \) from \( E_2 \) through \( E_7 \): \( E_2/E_1 = 4 \), \( E_3/E_2 = 2.25 \), \( E_4/E_3 = 1.78 \), \( E_5/E_4 = 1.56 \), \( E_6/E_5 = 1.44 \), \( E_7/E_6 = 1.36 \)

**Assess New Knowledge**

Have students use their calculated energy ratios from Demonstrate the Concept to make their own energy maps for hydrogen’s energy levels. Their energy maps will show clearly that hydrogen’s energy levels become more closely spaced as \( n \) increases.

---

**Portfolio**

**Models of the Atom**

- **Linguistic** Have students explain and trace the experimental evidence accompanying the evolution of models of the atom. Ask them to include Thomson’s plum-pudding model, Rutherford’s nuclear model, the Bohr model, and the quantum mechanical model. Have students place their explanations in their chemistry portfolios. [L2] [ELL] [C]

---

**Chemistry Journal**

**Orbital Shapes**

- **Visual-Spatial** Have students sketch the shapes and orientations of hydrogen’s 3s, 3p, and 3d orbitals. Have them label the orbital sketches and include them in their chemistry journals. [L2] [ELL] [C]

---

**Chemistry TEKS**

Pages 132–133

3(C), 3(E), 6(A)
Reteach

Explain that an electron’s position and velocity within an atomic orbital are not known. Reiterate that at a given instant, there is a 10% probability that the electron is outside the orbital’s 90% probability surface.

Extension

According to quantum mechanics, each electron in an atom can be described by four quantum numbers. Three of these ($n$, $l$, and $m_l$) are related to the probability of finding the electron at various points in space. The fourth ($m_s$) is related to the direction of electron spin—either clockwise or counterclockwise. The principal quantum number, $n$, specifies the atom’s energy level associated with the electron. $l$ specifies the energy sublevel and describes the shape of the region of space in which the electron moves. $m_l$ specifies the orientation in space of the orbital containing the electron. $m_s$ specifies the orientation of the electron’s spin axis.

Assessment

Knowledge  Ask students which hydrogen energy-level transition accounts for the violet line in its emission spectrum. $n = 6 \rightarrow n = 2$

Table 5-2 Hydrogen’s First Four Principal Energy Levels

<table>
<thead>
<tr>
<th>Principal quantum number $n$</th>
<th>Sublevels (types of orbitals) present</th>
<th>Number of orbitals related to sublevel</th>
<th>Total number of orbitals related to principal energy level $(n^2)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>s</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>s, p</td>
<td>3</td>
<td>4</td>
</tr>
<tr>
<td>3</td>
<td>s, p, d</td>
<td>5</td>
<td>9</td>
</tr>
<tr>
<td>4</td>
<td>s, p, d, f</td>
<td>7</td>
<td>16</td>
</tr>
</tbody>
</table>

Hydrogen’s first four principal energy levels, sublevels, and related atomic orbitals are summarized in Table 5-2. Note that the maximum number of orbitals related to each principal energy level equals $n^2$. Because each orbital may contain at most two electrons, the maximum number of electrons related to each principal energy level equals $2n^2$.

Given the fact that a hydrogen atom contains only one electron, you might wonder how the atom can have so many energy levels, sublevels, and related atomic orbitals. At any given time, the atom’s electron can occupy just one orbital. So you can think of the other orbitals as unoccupied spaces—spaces available should the atom’s energy increase or decrease. For example, when the hydrogen atom is in the ground state, the electron occupies the 1s orbital. However, the atom may gain a quantum of energy that excites the electron to the 2s orbital, to one of the three 2p orbitals, or to another vacant orbital.

You have learned a lot about electrons and quantized energy in this section: how Bohr’s orbits explained the hydrogen atom’s quantized energy states; how de Broglie’s insight led scientists to think of electrons as both particles and waves; and how Schrödinger’s wave equation predicted the existence of atomic orbitals containing electrons. In the next section, you’ll learn how the electrons are arranged in atomic orbitals of atoms having more than one electron.

Section 5.2 Assessment

13. According to the Bohr atomic model, why do atomic emission spectra contain only certain frequencies of light?
14. Why is the wavelength of a moving soccer ball not detectable to the naked eye?
15. What sublevels are contained in the hydrogen atom’s first four energy levels? What orbitals are related to each s sublevel and each p sublevel?
16. Thinking Critically  Use de Broglie’s wave-particle duality and the Heisenberg uncertainty principle to explain why the location of an electron in an atom is uncertain.
17. Comparing and Contrasting  Compare and contrast the Bohr model and quantum mechanical model of the atom.

The Heisenberg uncertainty principle states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Bohr model: the electron is a particle; the hydrogen atom has only certain allowable energy levels.
When you consider that atoms of the heaviest elements contain in excess of 100 electrons, that there are numerous principal energy levels and sublevels and their corresponding orbitals, and that each orbital may contain a maximum of two electrons, the idea of determining the arrangement of an atom’s electrons seems daunting. Fortunately, the arrangement of electrons in atoms follows a few very specific rules. In this section, you’ll learn these rules and their occasional exceptions.

Ground-State Electron Configurations

The arrangement of electrons in an atom is called the atom’s electron configuration. Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest possible energy. The most stable, lowest-energy arrangement of the electrons in atoms of each element is called the element’s ground-state electron configuration. Three rules, or principles—the aufbau principle, the Pauli exclusion principle, and Hund’s rule—define how electrons can be arranged in an atom’s orbitals.

The aufbau principle

The aufbau principle states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element’s ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in Figure 5-17. In the diagram, each box represents an atomic orbital. Several features of the aufbau diagram stand out.

- All orbitals related to an energy sublevel are of equal energy. For example, all three 2p orbitals are of equal energy.
- In a multi-electron atom, the energy sublevels within a principal energy level have different energies. For example, the three 2p orbitals are of higher energy than the 2s orbital.

![Fig. 5-17](Image)

The aufbau diagram shows the energy of each sublevel. Each box on the diagram represents an atomic orbital. Does the 3d or 4s sublevel have greater energy?

Vocabulary

- electron configuration
- aufbau principle
- Pauli exclusion principle
- Hund’s rule
- valence electron
- electron-dot structure

Focus Transparency

Before presenting the lesson, display Section Focus Transparency 19 on the overhead projector. Have students answer the accompanying questions using Section Focus Transparency Master 19.

Focus Transparency Questions

- Which seats in the arena are likely to be in more demand?
- Imagine that center court represents an atom’s nucleus. Which part of the arena represents energy levels? Which part represents individual orbitals?

Using Science Terms

Explain that the name aufbau is derived from the German aufbauen, which means “to build up.”
Demonstration

Emission Spectra

Purpose
To illustrate the relationship between the electron configurations of nonmetals and their emission spectra

Materials
Spectrum tubes (H and Ne); spectrum tube power supply; Flinn C-Spectra diffraction grating; colored pencils or chalk

Safety Precautions
Use care around the spectrum tube high voltage power supply. Spectrum tubes will get hot when used.

Procedure
An inexpensive alternative to a spectroscopic is to tape a small piece of the Flinn C-Spectra diffraction grating to a 3 × 5 inch card. Have students view the spectrum emitted from the lights in the classroom. Then, darken the room and have them view the excited neon atoms in the powered neon spectrum tubes. Use colored pencils to record the emission spectrum of neon as seen through their diffraction gratings. Remind students that neon contains 10 electrons. Now repeat the process using a hydrogen spectrum tube. Since hydrogen has 1 electron, ask...
Recall that the number of electrons in an atom equals the number of protons, which is designated by the element’s atomic number. Carbon, which has an atomic number of six, has six electrons in its configuration.

Another shorthand method for describing the arrangement of electrons in an element’s atoms is called electron configuration notation. This method designates the principal energy level and energy sublevel associated with each of the atom’s orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written \(1s^22s^22p^2\). Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in Table 5-3. To help you visualize the relative sizes and orientations of atomic orbitals, the filled 1s, 2s, 2p\(_x\), 2p\(_y\), and 2p\(_z\) orbitals of the neon atom are illustrated in Figure 5-18.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Orbital diagram</th>
<th>Electron configuration notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>1</td>
<td>[ ]</td>
<td>(1s^1)</td>
</tr>
<tr>
<td>Helium</td>
<td>2</td>
<td>[ ]</td>
<td>(1s^2)</td>
</tr>
<tr>
<td>Lithium</td>
<td>3</td>
<td>[ ] [ ]</td>
<td>(1s^22s^1)</td>
</tr>
<tr>
<td>Beryllium</td>
<td>4</td>
<td>[ ] [ ]</td>
<td>(1s^22s^2)</td>
</tr>
<tr>
<td>Boron</td>
<td>5</td>
<td>[ ] [ ] [ ]</td>
<td>(1s^22s^22p^1)</td>
</tr>
<tr>
<td>Carbon</td>
<td>6</td>
<td>[ ] [ ] [ ]</td>
<td>(1s^22s^22p^2)</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>7</td>
<td>[ ] [ ] [ ] [ ]</td>
<td>(1s^22s^22p^3)</td>
</tr>
<tr>
<td>Oxygen</td>
<td>8</td>
<td>[ ] [ ] [ ] [ ]</td>
<td>(1s^22s^22p^4)</td>
</tr>
<tr>
<td>Fluorine</td>
<td>9</td>
<td>[ ] [ ] [ ] [ ]</td>
<td>(1s^22s^22p^5)</td>
</tr>
<tr>
<td>Neon</td>
<td>10</td>
<td>[ ] [ ] [ ] [ ]</td>
<td>(1s^22s^22p^6)</td>
</tr>
</tbody>
</table>

Table 5-3

Electron Configurations and Orbital Diagrams for Elements in the First Two Periods

Figure 5-18 How many electrons does each of neon’s orbitals hold? Each orbital contains two electrons. How many electrons in total does neon’s electron cloud contain? ten electrons

Visual Learning

Table 5-3 Have students write an electron configuration notation that shows the orbital occupancy related to a phosphorus atom’s 3p sublevel. \(3s^13p^1\) A chlorine atom’s 3s and 3p sublevels. \(3s^23p^6\) Concept Development

Ask students to think about and explain the analogy between Hund’s rule and the behavior of total strangers as they board an empty bus. By and large, passengers sit in separate rows until people occupy all rows. Only when no more empty rows are available do two passengers occupy a single row. For electrons, the situation is much the same as they occupy orbitals related to a sublevel. Chemistry’s bus principle is known as Hund’s rule.

Figure 5-18 The 1s, 2s, and 2p orbitals of a neon atom overlap. How many electrons does each of neon’s orbitals hold? How many electrons in total does neon’s electron cloud contain?

Assessment

Skill Have students view the excited spectral tube of another element such as mercury. Ask them to predict if Hg will have more lines than neon and hydrogen because it has 80 electrons. No, Hg actually has fewer lines in the visible spectrum. However, there are many additional lines in mercury’s IR and UV spectra.

students to predict if there will be more or fewer lines in hydrogen’s spectrum.

Results

The red-orange spectrum of neon also contains some green lines. Usually only 3 of the 4 lines of hydrogen are visible.

Analysis

1. Write the electron configurations of neon and hydrogen. Ne: \(1s^22s^22p^6\), H: \(1s^1\)
2. What is the appearance of neon in the excited state? In the ground state, neon is a clear, colorless gas. In the excited state it gives off a red-orange light.
3. Of the two spectra viewed, did hydrogen or neon have more lines? Explain why. Neon has more lines than hydrogen because its ten electrons have a greater number of possible energy transitions.

5.3 Electron Configurations

5.3 Electron Configurations

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**Figure Caption Question**

**Figure 5-19** Which is filled first, the 5s or 4p orbital? The 4p orbital is filled first.

**Reinforcement**

Point out that some textbooks, reference books, and periodic tables show electron configurations written in energy-level sequence rather than in aufbau sequence. Reinforce that using the energy-level sequence for electron configurations does not render the aufbau sequence invalid.

Note that electron configuration notation usually does not show the orbital distributions of electrons related to a sublevel. It’s understood that a designation such as nitrogen’s 2p\(^3\) represents the orbital occupancy 2p\(_x\), 2p\(_y\), 2p\(_z\).

For sodium, the first ten electrons occupy 1s, 2s, and 2p orbitals. Then, according to the aufbau sequence, the eleventh electron occupies the 3s orbital. The electron configuration notation and orbital diagram for sodium are written

\[
\text{Na} \quad 1s^2 2s^2 2p^6 3s^1 \quad [\text{Ne}] 3s^1
\]

Noble-gas notation is a method of representing electron configurations of noble gases using bracketed symbols. For example, [He] represents the electron configuration for helium, 1s\(^2\), and [Ne] represents the electron configuration for neon, 1s\(^2\)2s\(^2\)2p\(^6\). Compare the electron configuration for neon with sodium’s configuration above. Note that the inner-level configuration for sodium is identical to the electron configuration for neon. Using noble-gas notation, sodium’s electron configuration can be shortened to the form [Ne]3s\(^1\). The electron configuration for an element can be represented using the noble-gas notation for the noble gas in the previous period and the electron configuration for the energy level being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in Table 5-4.

When writing electron configurations, you may refer to a convenient memory aid called a sublevel diagram, which is shown in Figure 5-19. Note that following the direction of the arrows in the sublevel diagram produces the sublevel sequence shown in the aufbau diagram of Figure 5-17.

**Exceptions to predicted configurations** You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, [Ar]4s\(^2\)3d\(^4\), and copper, [Ar]4s\(^2\)3d\(^9\), would prove to be incorrect. The correct configurations for these two elements are:

\[
\text{Cr} \quad [\text{Ar}]4s^13d^5 \quad \text{Cu} \quad [\text{Ar}]4s^13d^{10}
\]

The electron configurations for these two elements, as well as those of several elements in other periods, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

### Table 5-4

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Complete electron configuration</th>
<th>Electron configuration using noble-gas notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>11</td>
<td>1s(^2)2s(^2)2p(^6)3s(^1)</td>
<td>[Ne]3s(^1)</td>
</tr>
<tr>
<td>Magnesium</td>
<td>12</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)</td>
<td>[Ne]3s(^2)</td>
</tr>
<tr>
<td>Aluminum</td>
<td>13</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^1)</td>
<td>[Ne]3s(^2)3p(^1)</td>
</tr>
<tr>
<td>Silicon</td>
<td>14</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^2)</td>
<td>[Ne]3s(^2)3p(^2)</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>15</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^3)</td>
<td>[Ne]3s(^2)3p(^3)</td>
</tr>
<tr>
<td>Sulfur</td>
<td>16</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^4)</td>
<td>[Ne]3s(^2)3p(^4)</td>
</tr>
<tr>
<td>Chlorine</td>
<td>17</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^5)</td>
<td>[Ne]3s(^2)3p(^5)</td>
</tr>
<tr>
<td>Argon</td>
<td>18</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^6)</td>
<td>[Ne]3s(^2)3p(^6) or [Ar]</td>
</tr>
</tbody>
</table>

### The Development of Fireworks

Explain that the Chinese likely first used fireworks about the second century B.C. After inventing explosive black powder, which they called “gung pow,” the Chinese developed black-powder “crackers” that produced loud explosions. Most scholars believe that the Chinese used crackers to frighten off evil spirits and to celebrate weddings, births, battle victories, and eclipses of the Moon.

Fireworks became much more interesting and colorful in the 1830s, when Italian pyrotechnics experts added potassium chlorate to the mix. The potassium chlorate provided more oxygen for the chemical reaction, making it burn faster and hotter. This enabled the Italians to include various inorganic compounds that burn at high temperatures and create spectacular colors. Fireworks’ colors are due to energy-level transitions of electrons in the metal atoms of these inorganic compounds.
Writing Electron Configurations

Germanium (Ge), a semiconducting element, is commonly used in the manufacture of computer chips. What is the ground-state electron configuration for an atom of germanium?

**1. Analyze the Problem**

You are given the semiconducting element, germanium (Ge). Consult the periodic table to determine germanium’s atomic number, which also is equal to its number of electrons. Also note the atomic number of the noble gas element that precedes germanium in the table. Determine the number of additional electrons a germanium atom has compared to the nearest preceding noble gas, and then write out germanium’s electron configuration.

**2. Solve for the Unknown**

From the periodic table, germanium’s atomic number is determined to be 32. Thus, a germanium atom contains 32 electrons. The noble gas preceding germanium is argon (Ar), which has an atomic number of 18. Represent germanium’s first 18 electrons using the chemical symbol for argon written inside brackets.

[Ar]

The remaining 14 electrons of germanium’s configuration need to be written out. Because argon is a noble gas in the third period of the periodic table, it has completely filled 3s and 3p orbitals. Thus, the remaining 14 electrons fill the 4s, 3d, and 4p orbitals in order.

[Ar]4s²3d¹⁰4p²

Using the maximum number of electrons that can fill each orbital, write out the electron configuration.

[Ar]4s²3d¹⁰4p²

**3. Evaluate the Answer**

All 32 electrons in a germanium atom have been accounted for. The correct preceding noble gas (Ar) has been used in the notation, and the order of orbital filling for the fourth period is correct (4s, 3d, 4p).

---

**PRACTICE PROBLEMS**

18. Write ground-state electron configurations for the following elements.
   - a. bromine (Br)
   - b. strontium (Sr)
   - c. antimony (Sb)
   - d. rhenium (Re)
   - e. terbium (Tb)
   - f. titanium (Ti)

19. How many electrons are in orbitals related to the third energy level of a sulfur atom?

20. How many electrons occupy p orbitals in a chlorine atom?

21. What element has the following ground-state electron configuration? [Kr]5s²4d¹⁰5p³

22. What element has the following ground-state electron configuration? [Xe]6s²
Valence Electrons

Only certain electrons, called valence electrons, determine the chemical properties of an element. Valence electrons are defined as electrons in the atom’s outermost orbitals—generally those orbitals associated with the atom’s highest principal energy level. For example, a sulfur atom contains 16 electrons, only six of which occupy the outermost 3s and 3p orbitals, as shown by sulfur’s electron configuration. Sulfur has six valence electrons.

\[ \text{S} \left[ \text{Ne}\right]3s^23p^4 \]

Similarly, although a cesium atom contains 55 electrons, it has but one valence electron, the 6s electron shown in cesium’s electron configuration.

\[ \text{Cs} \left[ \text{Xe}\right]6s^1 \]

Francium, which belongs to the same group as cesium, also has a single valence electron.

\[ \text{Fr} \left[ \text{Rn}\right]7s^1 \]

Electron-dot structures

Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method. An atom’s electron-dot structure consists of the element’s symbol, which represents the atomic nucleus and inner-level electrons, surrounded by dots representing the atom’s valence electrons. The American chemist G. N. Lewis (1875–1946), devised the method while teaching a college chemistry class in 1902.

In writing an atom’s electron-dot structure, dots representing valence electrons are placed one at a time on the four sides of the symbol (they may be placed in any sequence) and then paired up until all are used. The ground-state electron configurations and electron-dot structures for the elements in the second period are shown in Table 5-5.

**Table 5-5**

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Electron configuration</th>
<th>Electron-dot structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>3</td>
<td>1s²2s¹</td>
<td>Li⁺</td>
</tr>
<tr>
<td>Beryllium</td>
<td>4</td>
<td>1s²2s²</td>
<td>Be⁺</td>
</tr>
<tr>
<td>Boron</td>
<td>5</td>
<td>1s²2s²2p¹</td>
<td>B⁺</td>
</tr>
<tr>
<td>Carbon</td>
<td>6</td>
<td>1s²2s²2p²</td>
<td>C⁺</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>7</td>
<td>1s²2s²2p³</td>
<td>N⁺</td>
</tr>
<tr>
<td>Oxygen</td>
<td>8</td>
<td>1s²2s²2p⁴</td>
<td>O⁺</td>
</tr>
<tr>
<td>Fluorine</td>
<td>9</td>
<td>1s²2s²2p⁵</td>
<td>F⁺</td>
</tr>
<tr>
<td>Neon</td>
<td>10</td>
<td>1s²2s²2p⁶</td>
<td>Ne⁺</td>
</tr>
</tbody>
</table>

**Chemistry Journal**

Another Solar System—What if?

**Linguistic** Ask students to write essays for their journals in which they speculate about flying a spacecraft to a planet in a different solar system. In the new solar system, they discover that each atomic orbital of the planet’s solid, liquid, and gaseous matter may contain up to three electrons rather than just two. Their speculation should focus on the characteristics of the elements on this new planet.
24. State the aufbau principle in your own words.

25. Apply the Pauli exclusion principle, the aufbau principle, and Hund’s rule to write out the electron configuration and draw the orbital diagram for each of the following elements.

   a. silicon
   b. fluorine
   c. calcium
   d. krypton


27. Thinking Critically Use Hund’s rule and orbital diagrams to describe the sequence in which ten electrons occupy the five orbitals related to an atom’s d sublevel.

28. Interpreting Scientific Illustrations Which of the following is the correct electron-dot structure for an atom of selenium? Explain.

   a. \( \text{Se} \)
   b. \( \text{Se} \)
   c. \( \text{Se} \)
   d. \( \text{Se} \)

Section 5.3 Assessment

24. Electrons tend to occupy the lowest-energy orbital available.

25. a. Si 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^2\)  
   \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   1s 2s 2p 3s 3p  

   b. F 1s\(^2\)2s\(^2\)2p\(^6\)  
   \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   1s 2s 2p  

26. A valence electron is an electron in an atom’s outermost orbitals.

   a. \( \cdot \text{Si} \cdot \)  
   c. \( \cdot \text{Ca} \cdot \)  

   b. \( \cdot \text{F} \cdot \)  
   d. \( \cdot \text{K} \cdot \)  

27. 1 electron \( \uparrow \)  
   2 electrons \( \uparrow \uparrow \)  
   3 electrons \( \uparrow \uparrow \uparrow \)  
   4 electrons \( \uparrow \uparrow \uparrow \uparrow \)  
   5 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   6 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   7 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   8 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   9 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  
   10 electrons \( \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow \)  

   Single electrons with the same spin occupy each equal-energy orbital before additional electrons with opposite spins occupy the same orbital.

28. c is correct; a shows three two-electron orbitals; b shows one three-electron orbital; d has the wrong symbol.
Pre-Lab

1. Read the entire CHEMLAB.
2. Explain how electrons in an element’s atoms produce an emission spectrum.
3. Distinguish among a continuous spectrum, an emission spectrum, and an absorption spectrum.
4. Prepare your data tables.

Procedure

1. Use a Flinn C-Spectra® to view an incandescent light bulb. What do you observe? Draw the spectrum using colored pencils.

Expected Results

For each colored solution listed below, all colors are visible except as noted.

- Red solution: blue and green
- Green solution: red and orange
- Blue solution: yellow, orange, and some red
- Yellow solution: blue

Materials

(For each group)
- ring stand with clamp
- 40-W tubular light bulb
- light socket with power cord
- 275-mL polystyrene culture flask (4)
- Flinn C-Spectra® or similar diffraction grating
- food coloring (red, green, blue, and yellow)
- set of colored pencils
- book

(For entire class)
- spectrum tubes (hydrogen, neon, and mercury)
- spectrum tube power supplies (3)


drawings of absorption spectra

<table>
<thead>
<tr>
<th>Red</th>
<th>Green</th>
<th>Blue</th>
<th>Yellow</th>
</tr>
</thead>
</table>

drawings of emission spectra

<table>
<thead>
<tr>
<th>Hydrogen</th>
<th>Neon</th>
<th>Mercury</th>
</tr>
</thead>
</table>

Line Spectra

You know that sunlight is made up of a continuous spectrum of colors that combine to form white light. You also have learned that atoms of gases can emit visible light of characteristic wavelengths when excited by electricity. The color you see is the sum of all of the emitted wavelengths. In this experiment, you will use a diffraction grating to separate these wavelengths into emission line spectra.

You also will investigate another type of line spectrum—the absorption spectrum. The color of each solution you observe is due to the reflection or transmission of unabsorbed wavelengths of light. When white light passes through a sample and then a diffraction grating, dark lines show up on the continuous spectrum of white light. These lines correspond to the wavelengths of the photons absorbed by the solution.

Problem

What absorption and emission spectra do various substances produce?

Objectives

- Observe emission spectra of several gases.
- Observe the absorption spectra of various solutions.
- Analyze patterns of absorption and emission spectra.

Safety Precautions

- Always wear safety goggles and a lab apron.
- Use care around the spectrum tube power supplies.
- Spectrum tubes will get hot when used.

Disposal

You may want to reuse the flasks of food coloring solutions.

Preparation of Materials

- Set up light sockets with light bulbs prior to class and have them plugged in.
- Set up spectrum power supplies and tubes prior to class.

Pre-Lab

- When electrons drop from higher-energy orbitals to lower-energy orbitals, the atom emits energy in the form of light. Each orbital transition is associated with a characteristic spectral line.
- A continuous spectrum contains a continuum of visible colors from red to violet. An absorption spectrum is a continuous spectrum containing black lines at wavelengths associated with the atoms’ energy absorptions. An emission spectrum consists of colored lines associated with the atoms’ energy-level transitions.
2. Use the Flinn C-Spectra® to view the emission spectra from tubes of gaseous hydrogen, neon, and mercury. Use colored pencils to make drawings in the data table of the spectra observed.

3. Fill a 275-mL culture flask with about 100-mL water. Add 2 or 3 drops of red food coloring to the water. Shake the solution.

4. Repeat step 3 for the green, blue, and yellow food coloring. CAUTION: Be sure to thoroughly dry your hands before handling electrical equipment.

5. Set up the light 40-W light bulb so that it is near eye level. Place the flask with red food coloring about 8 cm from the light bulb. Use a book or some other object to act as a stage to put the flask on. You should be able to see light from the bulb above the solution and light from the bulb projecting through the solution.

6. With the room lights darkened, view the light using the Flinn C-Spectra®. The top spectrum viewed will be a continuous spectrum of the white light bulb. The bottom spectrum will be the absorption spectrum of the red solution. The black areas of the absorption spectrum represent the colors absorbed by the red food coloring in the solution. Use colored pencils to make a drawing in the data table of the absorption spectra you observed.

7. Repeat steps 5 and 6 using the green, blue, and yellow colored solutions.

**Cleanup and Disposal**

1. Turn off the light socket and spectrum tube power supplies.
2. Wait several minutes to allow the incandescent light bulb and the spectrum tubes to cool.
3. Follow your teacher’s instructions on how to dispose of the liquids and how to store the light bulb and spectrum tubes.

**Analyze and Conclude**

1. Thinking Critically How can the existence of spectra help to prove that energy levels in atoms exist?
2. Thinking Critically How can the single electron in a hydrogen atom produce all of the lines found in its emission spectrum?
3. Predicting How can you predict the absorption spectrum of a solution by looking at its color?
4. Thinking Critically How can spectra be used to identify the presence of specific elements in a substance?

**Real-World Chemistry**

1. How can absorption and emission spectra be used by the Hubble space telescope to study the structures of stars or other objects found in deep space?
2. The absorption spectrum of chlorophyll a indicates strong absorption of red and blue wavelengths. Explain why leaves appear green.

**The approximate emission spectra of the gas spectrum tubes are shown below.**

<table>
<thead>
<tr>
<th>Gas</th>
<th>Spectrum</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>Orange</td>
</tr>
<tr>
<td>Mercury</td>
<td>Blue</td>
</tr>
<tr>
<td>Neon</td>
<td>Yellow</td>
</tr>
</tbody>
</table>

**Analyze and Conclude**

1. The spectral lines indicate energy is absorbed or released as the atom transitions from one energy level to another.
2. At any given time, the electron occupies a single orbital. However, it can move into other, vacant orbitals as the atom absorbs or emits energy.
3. The color of a solution is due to the color of light it transmits. The colors not transmitted are absorbed, and these colors comprise the absorption spectrum.
4. The spectrum of each element is unique. Thus, the presence of a unique atomic spectrum indicates the presence of that element.

**Real-World Chemistry**

1. The light emitted by stars can be analyzed for the presence of unique atomic spectra. Such spectra can identify the types of matter that comprise the star.
2. Leaves appear green because they reflect (do not absorb) green light. The reflected green light is what our eyes see.
How It Works

Lasers

A laser is a device that produces a beam of intense light of a specific wavelength (color). Unlike light from a flashlight, laser light is coherent; that is, it does not spread out as it travels through space. The precise nature of lasers led to their use in pointing and aiming devices, CD players, optical fiber data transmission, and surgery.

1. The spiral-wound high-intensity lamp flashes, supplying energy to the helium-neon gas mixture inside the tube. The atoms of the gas absorb the light energy and are raised to an excited energy state.

2. The excited atoms begin returning to the ground state, emitting photons in the process. These initial photons travel in all directions.

3. The emitted photons hit other excited atoms, causing them to release additional photons. These additional photons are the same wavelength as the photons that struck the excited atoms, and they are coherent (their waves are in sync because they are identical in wavelength and direction).

4. Photons traveling parallel to the tube are reflected back through the tube by the flat mirrors located at each end. The photons strike additional excited atoms and cause more photons to be released. The intensity of the light in the tube builds.

5. Some of the laser’s coherent light passes through the partially transparent mirror at one end of the tube and exits the laser. These photons make up the light emitted by the laser.

Thinking Critically

1. Inferring How does the material used in the laser affect the type of light emitted?

2. Relating Cause and Effect Why is one mirror partially transparent?

Thinking Critically

1. The arrangement of electrons varies from one substance to another. As a result, the characteristics of light emitted by the laser also vary. The material in the laser determines the characteristics of the light produced.

2. If one of the mirrors were not partially transparent, the photons would have no way to escape the laser. If the mirror were totally transparent, the photons would exit the laser after just one pass through the tube. Thus, the photons could not stimulate the emission of additional photons; the laser would soon weaken.
Summary

5.1 Light and Quantized Energy
- All waves can be described by their wavelength, frequency, amplitude, and speed.
- Light is an electromagnetic wave. In a vacuum, all electromagnetic waves travel at a speed of $3.00 \times 10^8$ m/s.
- All electromagnetic waves may be described as both waves and particles. Particles of light are called photons.
- Energy is emitted and absorbed by matter in quanta.
- In contrast to the continuous spectrum produced by white light, an element's atomic emission spectrum consists of a series of fine lines of individual colors.

5.2 Quantum Theory and the Atom
- According to the Bohr model of the atom, hydrogen's atomic emission spectrum results from electrons dropping from higher-energy atomic orbits to lower-energy atomic orbits.
- The de Broglie equation predicts that all moving particles have wave characteristics and relates each particle's wavelength to its mass, its frequency, and Planck's constant.

Key Equations and Relationships
- EM Wave relationship: $c = \lambda \nu$ (p. 119)
- Energy of a quantum: $E_{\text{quantum}} = h \nu$ (p. 123)
- Energy of a photon: $E_{\text{photon}} = h \nu$ (p. 124)
- Energy change of an electron: $\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}}$ (p. 128)
- $\Delta E = E_{\text{photon}} = h \nu$ (p. 128)
- de Broglie's equation: $\lambda = \frac{h}{mv}$ (p. 130)

Vocabulary
- amplitude (p. 119)
- atomic emission spectrum (p. 120)
- atomic orbital (p. 132)
- aufbau principle (p. 135)
- de Broglie equation (p. 130)
- electromagnetic radiation (p. 118)
- electromagnetic spectrum (p. 120)
- electron configuration (p. 135)
- electron-dot structure (p. 140)
- energy sublevel (p. 133)
- frequency (p. 118)
- ground state (p. 127)
- Heisenberg uncertainty principle (p. 131)
- Hund's rule (p. 136)
- Pauli exclusion principle (p. 136)
- photoelectric effect (p. 123)
- photon (p. 123)
- Planck's constant (p. 123)
- principal energy level (p. 133)
- principal quantum number (p. 132)
- quantum (p. 122)
- quantum mechanical model of the atom (p. 131)
- valence electron (p. 140)
- wavelength (p. 118)

Portfolio Options
Review the portfolio options that are provided throughout the chapter. Encourage students to select one product that demonstrates their best work for the chapter. Have students explain what they have learned and why they chose this example for placement into their portfolios. Additional portfolio options may be found in the Challenge Problems booklet of the Teacher Classroom Resources.

Using the Vocabulary
To reinforce chapter vocabulary, have students write a sentence using each term.

Review Strategies
- Have students describe the electromagnetic spectrum and differentiate between visible light and infrared radiation.
- Ask students to write the equation that relates an electromagnetic wave's frequency and wavelength.
- Have students write the equation that relates the energy of a quantum of electromagnetic radiation to the frequency of an associated wave.
- Ask students to explain the significance of Heisenberg's uncertainty principle as it relates to electrons in atoms.
- Have students explain the relationship between an atom's orbitals and its energy levels.
- Problems from Appendix A or the Supplemental Problems booklet can be used for review.

The Princeton Review
Reviewing Chemistry is a component of the Teacher Classroom Resources package that was prepared by The Princeton Review. Use the Chapter 5 review materials in this book to review the chapter with your students.
Concept Mapping

29. Complete the concept map using the following terms:
   speed, \( c = \lambda \nu \); electromagnetic waves, wavelength, characteristic properties, frequency, \( c \), and hertz.

Mastering Concepts

30. a. Frequency is the number of waves that pass a given point per second.
    b. Wavelength is the shortest distance between equivalent points on a continuous wave.
    c. A quantum is the minimum amount of energy that can be lost or gained by an atom.
    d. An atom’s ground state is its lowest allowable energy state.

31. Typical answers will say that the model did not explain the following: how the atom’s negatively charged electrons occupy the space around the nucleus; why the electrons are not drawn into the atom’s positively charged nucleus; a rationale for the chemical properties of the elements.

32. light, microwaves, X rays, radio waves

33. Electricity passing through the tube excites neon atoms to higher energy levels. As the excited atoms drop back to lower energy levels, they emit light.

34. a particle of electromagnetic radiation having a rest mass of zero and carrying a quantum of energy

35. a phenomenon in which a metal emits electrons when light of a sufficient frequency shines on it

36. According to Planck, for a given frequency, \( \nu \), matter can emit or absorb energy only in discrete quanta that are whole-number multiples of \( h \nu \).

37. He proposed that photons must have a certain minimum, or threshold, value to cause the ejection of a photoelectron.

38. d. X rays, a. ultraviolet light, b. microwaves, c. radio waves

40. According to the Bohr model, how do electrons move in atoms? (5.2)

41. What does \( n \) designate in Bohr’s atomic model? (5.2)

42. Why are you unaware of the wavelengths of moving objects such as automobiles and tennis balls? (5.2)

43. What is the name of the atomic model in which electrons are treated as waves? Who first wrote the electron wave equations that led to this model? (5.2)

44. What is an atomic orbital? (5.2)

45. What is the probability that an electron will be found within an atomic orbital? (5.2)

46. What does \( n \) represent in the quantum mechanical model of the atom? (5.2)

47. How many energy sublevels are contained in each of the hydrogen atom’s first three energy levels? (5.2)

48. What atomic orbitals are related to a p sublevel? To a d sublevel? (5.2)

49. Which of the following atomic orbital designations are incorrect? (5.2)
   a. 7f
   b. 3f
   c. 2d
   d. 6p

50. What do the sublevel designations s, p, d, and f specify with respect to the atom’s orbitals? (5.2)

51. What do subscripts such as \( y \) and \( xz \) tell you about atomic orbitals? (5.2)

52. What is the maximum number of electrons an orbital may contain? (5.2)

53. Why is it impossible to know precisely the velocity and position of an electron at the same time? (5.2)

54. What shortcomings caused scientists to finally reject Bohr’s model of the atom? (5.2)

55. Describe de Broglie’s revolutionary concept involving the characteristics of moving particles. (5.2)

56. How is an orbital’s principal quantum number related to the atom’s major energy levels? (5.2)

57. Explain the meaning of the aufbau principle as it applies to atoms with many electrons. (5.3)

58. In what sequence do electrons fill the atomic orbitals related to a sublevel? (5.3)

59. Why must the two arrows within a single block of an orbital diagram be written in opposite (up and down) directions? (5.3)

60. How does noble-gas notation shorten the process of writing an element’s electron configuration? (5.3)

61. What are valence electrons? How many of a magnesium atom’s 12 electrons are valence electrons? (5.3)
62. Light is said to have a dual wave-particle nature. What does this statement mean? (5.3)
63. Describe the difference between a quantum and a photon. (5.3)
64. How many electrons are shown in the electron-dot structures of the following elements? (5.3)
   a. carbon    c. calcium
   b. iodine    d. gallium

Mastering Problems
Wavelength, Frequency, Speed, and Energy (5.1)
65. What is the wavelength of electromagnetic radiation having a frequency of 5.00 \( \times \) 10^{12} Hz? What kind of electromagnetic radiation is this? (5.3)
66. What is the frequency of electromagnetic radiation having a wavelength of 3.33 \( \times \) 10^{-8} m? What type of electromagnetic radiation is this? (5.3)
67. The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light? (5.3)
68. What is the speed of an electromagnetic wave having a frequency of 1.33 \( \times \) 10^{17} Hz and a wavelength of 2.25 nm? (5.3)
69. Use Figure 5-5 to determine each of the following types of radiation.
   a. radiation with a frequency of 8.6 \( \times \) 10^{11} s^{-1}
   b. radiation with a wavelength of 4.2 nm
   c. radiation with a frequency of 5.6 MHz
   d. radiation that travels at a speed of 3.00 \( \times \) 10^{8} m/s
70. What is the energy of a photon of red light having a frequency of 4.48 \( \times \) 10^{14} Hz? (5.3)
71. Mercury’s atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom? (5.3)

72. What is the energy of an ultraviolet photon having a wavelength of 1.18 \( \times \) 10^{-8} m? (5.3)
73. A photon has an energy of 2.93 \( \times \) 10^{-25} J. What is its frequency? What type of electromagnetic radiation is the photon? (5.3)

74. A photon has an energy of 1.10 \( \times \) 10^{-13} J. What is the photon’s wavelength? What type of electromagnetic radiation is it? (5.3)
75. How long does it take a radio signal from the Voyager spacecraft to reach Earth if the distance between Voyager and Earth is 2.72 \( \times \) 10^{10} km? (5.3)
76. If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station’s signal in meters? What is the energy of a photon of the station’s electromagnetic signal? (5.3)

Electron Configurations (5.3)
77. List the aufbau sequence of orbitals from 1s to 7p. (5.3)
78. Write orbital notations and complete electron configurations for atoms of the following elements.
   a. beryllium
   b. aluminum
   c. nitrogen
   d. sodium
79. Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.
   a. Mn
   b. Kr
   c. P
   d. Zn
   e. Zr
70. What elements are represented by each of the following electron configurations?
   a. 1s^22s^22p^3
   b. [Ar]4s^2
   c. [Xe]6s^24f^4
   d. [Kr]5s^24d^{10}5p^4
   e. [Rn]7s^25f^{13}
   f. 1s^22s^22p^33s^23p^44s^3d^{10}5p^3
81. Draw electron-dot structures for atoms of each of the following elements.
   a. carbon
   b. arsenic
   c. polonium
   d. potassium
   e. barium
82. An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom’s n = 4 principal energy level? (5.3)

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56. Because the orbital’s principal quantum number indicates the orbital’s relative size and energy, it also specifies the atom’s major energy level. (5.3)
57. The aufbau principle describes the sequence in which an atom’s orbitals are filled with electrons. (5.3)
58. Each orbital must contain a single electron before any orbital contains two electrons. (5.3)
59. Two electrons occupying a single atomic orbital must have opposite spins. (5.3)
60. The noble-gas notation uses the bracketed symbol of the preceding noble gas in the periodic table to represent an atom’s inner electrons. (5.3)
61. Valence electrons are the electrons in an atom’s outermost orbitals; 2 (5.3)
62. Light exhibits wavelike behavior in some situations and particelleike behavior in others. (5.3)
63. Each orbital must contain a single electron before any orbital contains two electrons. (5.3)
64. How many electrons are shown in the electron-dot structures of the following elements? (5.3)
   a. carbon    c. calcium
   b. iodine    d. gallium
65. What is the wavelength of electromagnetic radiation having a frequency of 5.00 \( \times \) 10^{12} Hz? What kind of electromagnetic radiation is this? (5.3)
66. What is the frequency of electromagnetic radiation having a wavelength of 3.33 \( \times \) 10^{-8} m? What type of electromagnetic radiation is this? (5.3)
67. The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light? (5.3)
68. What is the speed of an electromagnetic wave having a frequency of 1.33 \( \times \) 10^{17} Hz and a wavelength of 2.25 nm? (5.3)
69. Use Figure 5-5 to determine each of the following types of radiation.
   a. radiation with a frequency of 8.6 \( \times \) 10^{11} s^{-1}
   b. radiation with a wavelength of 4.2 nm
   c. radiation with a frequency of 5.6 MHz
   d. radiation that travels at a speed of 3.00 \( \times \) 10^{8} m/s
70. What is the energy of a photon of red light having a frequency of 4.48 \( \times \) 10^{14} Hz? (5.3)
71. Mercury’s atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom? (5.3)

72. What is the energy of an ultraviolet photon having a wavelength of 1.18 \( \times \) 10^{-8} m? (5.3)
73. A photon has an energy of 2.93 \( \times \) 10^{-25} J. What is its frequency? What type of electromagnetic radiation is the photon? (5.3)
74. A photon has an energy of 1.10 \( \times \) 10^{-13} J. What is the photon’s wavelength? What type of electromagnetic radiation is it? (5.3)
75. How long does it take a radio signal from the Voyager spacecraft to reach Earth if the distance between Voyager and Earth is 2.72 \( \times \) 10^{10} km? (5.3)
76. If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station’s signal in meters? What is the energy of a photon of the station’s electromagnetic signal? (5.3)
77. List the aufbau sequence of orbitals from 1s to 7p. (5.3)
78. Write orbital notations and complete electron configurations for atoms of the following elements.
   a. beryllium
   b. aluminum
   c. nitrogen
   d. sodium
79. Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.
   a. Mn
   b. Kr
   c. P
   d. Zn
   e. Zr
80. What elements are represented by each of the following electron configurations?
   a. 1s^22s^22p^3
   b. [Ar]4s^2
   c. [Xe]6s^24f^4
   d. [Kr]5s^24d^{10}5p^4
   e. [Rn]7s^25f^{13}
   f. 1s^22s^22p^33s^23p^44s^3d^{10}5p^3
81. Draw electron-dot structures for atoms of each of the following elements.
   a. carbon
   b. arsenic
   c. polonium
   d. potassium
   e. barium
82. An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom’s n = 4 principal energy level? (5.3)
**CHAPTER 5 ASSESSMENT**

63. A quantum is the minimum amount of energy that can be lost or gained by an atom, while a photon is a particle of light that carries a quantum of energy.

64. **Mastering Problems**
   Complete solutions to Chapter Assessment problems can be found in the Solutions Manual.

Wavelength, Frequency, Speed, and Energy (5.1)

**Level 1**

65. \( \lambda = 6.00 \times 10^{-5} \text{ m}; \) infrared radiation
66. \( \nu = 9.01 \times 10^{15} \text{ s}^{-1}; \) ultraviolet radiation
67. \( \nu = 3.8 \times 10^{14} \text{ s}^{-1}; \) FM wave
68. \( \nu = 3.00 \times 10^8 \text{ m/s}; \) AM radio
69. a. infrared
   b. X ray
   c. AM radio
   d. any EM wave

**Level 2**

70. \( E_{\text{photon}} = 2.97 \times 10^{-19} \text{ J} \)
71. \( \lambda = 615 \text{ nm}, 4.88 \times 10^{14} \text{ s}^{-1}, E_{\text{photon}} = 3.23 \times 10^{-19} \text{ J} \)
72. \( E_{\text{photon}} = 1.68 \times 10^{-17} \text{ J} \)
73. \( \nu = 4.42 \times 10^{13} \text{ s}^{-1}; \) TV or FM wave
74. \( \lambda = 1.81 \times 10^{-12} \text{ m}, \) an X ray or gamma ray
75. \( t = 9070 \text{ s}, \) or 151 min
76. \( \lambda = 2.87 \text{ m}, E_{\text{photon}} = 6.92 \times 10^{-26} \text{ J} \)

**Electron Configurations (5.3)**

**Level 1**

77. 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p

78. a. Be 1s²2s²
   b. Al 1s²2s²2p⁶3s³3p³
   c. N 1s²2s²2p³
   d. Na 1s²2s²2p⁶3s¹

**Level 2**

79. a. Mn [Ar]4s²3d⁵
   b. Kr [Ar]4s³3d¹⁰4p⁶
   c. P [Ne]3s²3p³
   d. Zn [Ar]4s²3d¹⁰
   e. Zr [Kr]5s²⁴d²
   f. W [Xe]6s²4f¹⁴5d⁴
   g. Pb [Xe]6s²4f¹⁴5d¹⁰6p²
   h. Ra [Rn]7s²
   i. Sm [Xe]6s²4f⁶
   j. Bk [Rn]7s²5f³

**Thinking Critically**

96. **Comparing and Contrasting** Briefly discuss the difference between an orbit in Bohr’s model of the atom and an orbital in the quantum mechanical view of the atom.

97. **Applying Concepts** Scientists use atomic emission spectra to determine the elements in materials of unknown composition. Explain what makes this method possible.

98. **Using Numbers** It takes \( 8.17 \times 10^{-19} \text{ J} \) of energy to remove one electron from a gold surface. What is the maximum wavelength of light capable of causing this effect?

99. **Drawing a Conclusion** The elements aluminum, silicon, gallium, germanium, arsenic, selenium are all used in making various types of semiconductor devices. Write electron configurations and electron-don structures for atoms of each of these elements. What similarities among the elements’ electron configurations do you notice?

**Writing in Chemistry**

100. In order to make “neon” signs emit a variety of colors, manufacturers often fill the signs with gases other than neon. Research the use of gases in neon signs and specify the colors produced by the gases.

**Cumulative Review**

Refresh your understanding of previous chapters by answering the following.

101. Round 20.561 20 g to three significant figures. (Chapter 2)

102. Identify each of the following as either chemical or physical properties of the substance. (Chapter 3)
   a. mercury is a liquid at room temperature
   b. sucrose is a white, crystalline solid
   c. iron rusts when exposed to moist air
   d. paper burns when ignited

103. Identify each of the following as a pure substance or a mixture. (Chapter 3)
   a. distilled water
   b. orange juice with pulp
   c. milk
   d. diamond
   e. milk
   f. copper metal

104. An atom of gadolinium has an atomic number of 64 and a mass number of 153. How many electrons, protons, and neutrons does it contain? (Chapter 4)

148 Chapter 5 Electrons in Atoms
Use these questions and the text-taking tip to prepare for your standardized test.

1. Cosmic rays are high-energy radiation from outer space. What is the frequency of a cosmic ray that has a wavelength of $2.67 \times 10^{-13}$ m when it reaches Earth? (The speed of light is $3.00 \times 10^8$ m/s.)
   - a. $8.90 \times 10^{-22}$ s⁻¹
   - b. $3.75 \times 10^{12}$ s⁻¹
   - c. $8.01 \times 10^{-9}$ s⁻¹
   - d. $1.12 \times 10^{15}$ s⁻¹

2. Wavelengths of light between $5.75 \times 10^{-9}$ m and $5.85 \times 10^{-9}$ m appear yellow to the human eye. What is the energy of a photon of yellow light having a frequency of $5.45 \times 10^{16}$ s⁻¹? (Planck’s constant is $6.626 \times 10^{-34}$ J·s.)
   - a. $3.61 \times 10^{-17}$ J
   - b. $1.22 \times 10^{-16}$ J
   - c. $8.23 \times 10^{-16}$ J
   - d. $3.81 \times 10^{-16}$ J

3. Using noble-gas notation, the ground-state electron configuration of Cd is __________.
   - a. [Kr]4d⁰4f⁰
   - b. [Ar]4s²3d⁶
   - c. [Kr]5s²4d⁴
   - d. [Xe]6s²4f⁴5d⁸

4. The element that has the ground-state electron configuration [Xe]6s²4f⁴5d⁸ is __________.
   - a. La
dc. W
   - b. Ti
d. Os

5. The complete electron configuration of a scandium atom is __________.
   - a. 1s²2s²2p⁶3s²3p⁶4s²3d⁵
   - b. 1s²2s²2p⁶3s²3p⁶4s²3d⁵
   - c. 1s²2s²2p⁶3s²3p⁶4s²3d⁵
   - d. 1s²2s²2p⁶3s²3p⁶4s²3d⁵

6. Which of the following is the correct orbital diagram for the third and fourth principal energy levels of vanadium?
   - a. [Ar]4s²3d²
   - b. [Ar]4s²3d³
   - c. [Ar]4s²3d⁴
   - d. [Ar]4s²3d⁵

7. Which of the following orbitals has the highest energy?
   - a. 4f
   - b. 5p
   - c. 6s
   - d. 3d

8. What is the electron-dot structure for iodine?
   - a. \( \cdot I \cdot \)
   - b. \( \cdot I \cdot \)
   - c. \( \cdot I \cdot \)
   - d. \( \cdot I \cdot \)

9. The picture below shows all of the orbitals related to one type of sublevel. The type of sublevel to which these orbitals belong is __________.

10. What is the maximum number of electrons related to the fifth principal energy level of an atom?
    - a. 10
    - b. 20
    - c. 25
    - d. 50

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**Standardized Test Practice**

**Cumulative Review**

101. 20.6 g

102. a. physical property
     b. physical property
     c. chemical property
     d. chemical property

103. a. pure substance
     b. mixture
     c. mixture

**Mixed Review**

83. \( \nu = 3.00 \times 10^8 \text{ s}^{-1} \)
84. a. 18   c. 72
     b. 32   d. 98
85. \( \lambda = 5.20 \times 10^{-7} \text{ m} \)
86. a. longest wavelength: 4
     b. greatest frequency: 3
     c. largest amplitude: 1
     d. shortest wavelength: 3
87. a. 1   c. 5
     b. 3   d. 7
88. 1s²2s²2p⁶3s²3p⁶4s²3d⁸
     a. [Ar]4s²3d⁸
89. helium, calcium, cobalt, barium
90. \( n = 4 \rightarrow n = 2 \)
91. The two dots are the atom’s two 4s valence electrons.
92. \( \nu = 4.54 \times 10^{15} \text{ s}^{-1} \); \( \lambda = 6.60 \times 10^{-8} \text{ m} \)
93. boron
94. francium
95. \( 3.10 \times 10^{19} \text{ photons} \)

**Thinking Critically**

96. In the Bohr model, an orbit is a circular path taken by an electron as it moves around the atomic nucleus. In the quantum mechanical model, an orbital is a three-dimensional region around the nucleus that describes the electron’s probable location.
97. Each element emits a characteristic, unique atomic emission spectrum.
98. \( \lambda = 2.43 \times 10^{-7} \text{ m} \)
99. Al [Ne]3s³3p¹
    Si [Ne]3s³3p²
    Ga [Ar]4s²3d¹0p¹
    Ge [Ar]4s²3d¹0p²
    As [Ar]4s²3d¹0p³
    Se [Ar]4s²3d¹0p⁴
    The atoms have filled s orbitals and p orbitals containing 1 to 4 electrons. See the Solutions Manual for electron dot structures.

**Do Some Reconnaissance** Find out what the conditions will be for taking the test. Is it timed or untimed? Can you eat a snack at the break? Can you use a calculator or other tools? Will those tools be provided? Will mathematical constants be given? Know these things in advance so that you can practice taking tests under the same conditions.